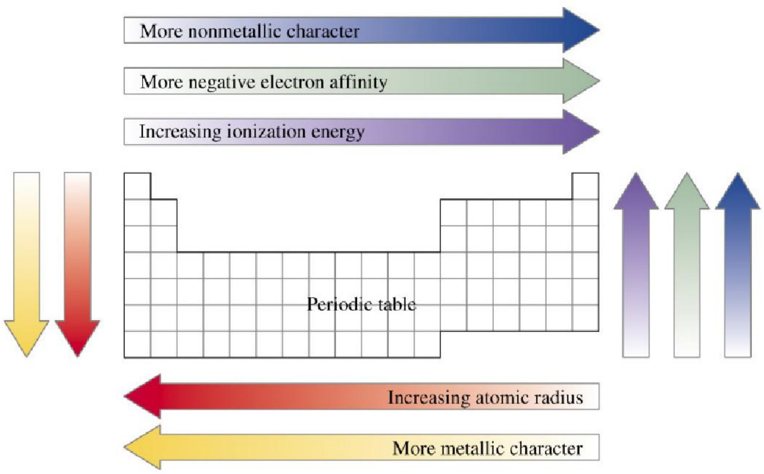
PERIODIC CLASSIFICATION

Table of Contents

- Theory 2
- Solved Examples 12
- Exercise - 1 : Basic Objective Questions 22
- Exercise - 2 : Previous Year JEE Mains Questions 25
- Exercise - 3 : Advanced Objective Questions 28
- Exercise - 4 : Previous Year JEE Advanced Questions 33
- Answer Key 35



THEORY

1. INTRODUCTION

Periodic table may be defined as the table which classifies all the known elements in accordance with their properties in such a way that elements with similar properties are grouped together in the same vertical column and dissimilar elements are separated from one another.

2. HISTORICAL DEVELOPMENT OF THE PERIODIC TABLE

All earlier attempts of the classification of the elements were based upon their atomic weights.

2.1 Dobereiner's Triads

In 1829, Dobereiner classified certain elements in the groups of three called **triads**. The three elements in a triad had similar chemical properties. When the elements in a triad were arranged in the order of increasing atomic weights, the atomic weight of the middle element was found to be approximately equal to the arithmetic mean of the other two elements.

1. Triad	**Iron**	**Cobalt**	**Nickel**	**Mean of 1st and 3rd**
At. wt.	55.85	58.93	58.71	Atomic weights are nearly the same
2. Triad	**Lithium**	**Sodium**	**Potassium**	
At. wt.	7	23	39	$\frac{7+39}{2}=23$
3. Triad	**Chlorine**	**Bromine**	**Iodine**	
At. wt.	35.5	80	127	$\frac{35.5+127}{2}=81.25$
4. Triad	**Calcium**	**Strontium**	**Barium**	
At. wt.	40	87.5	137	$\frac{40+177}{2}=88.5$

2.2 Newland's Law of Octaves

In 1865, an English chemist, *John Alexander Newlands* observed that

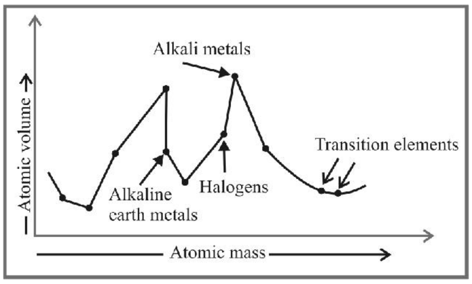
When the lighter elements were arranged in order of their increasing atomic weights, the properties of every eighth element were similar to those of the first one like the eighth note of a musical scale. This generalization was named as ***Newlands's law of octaves.***

Element	**Li**	**Be**	**B**	**C**	**N**	**O**	**F**
At. wt.	7	9	11	12	14	16	19
Element	**Na**	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**
At. wt.	23	24	27	29	31	32	35.5
Element	**K**	**Ca**					
At. wt.	39	40					

2.3 Lothar Meyer's Curve

"Physical properties of elements are periodic functions of their atomic masses."

According to Lothar Meyer, elements having similar properties occupy the similar positions in atomic volume viz atomic mass curve



2.4 Mendeleev's Periodic Law

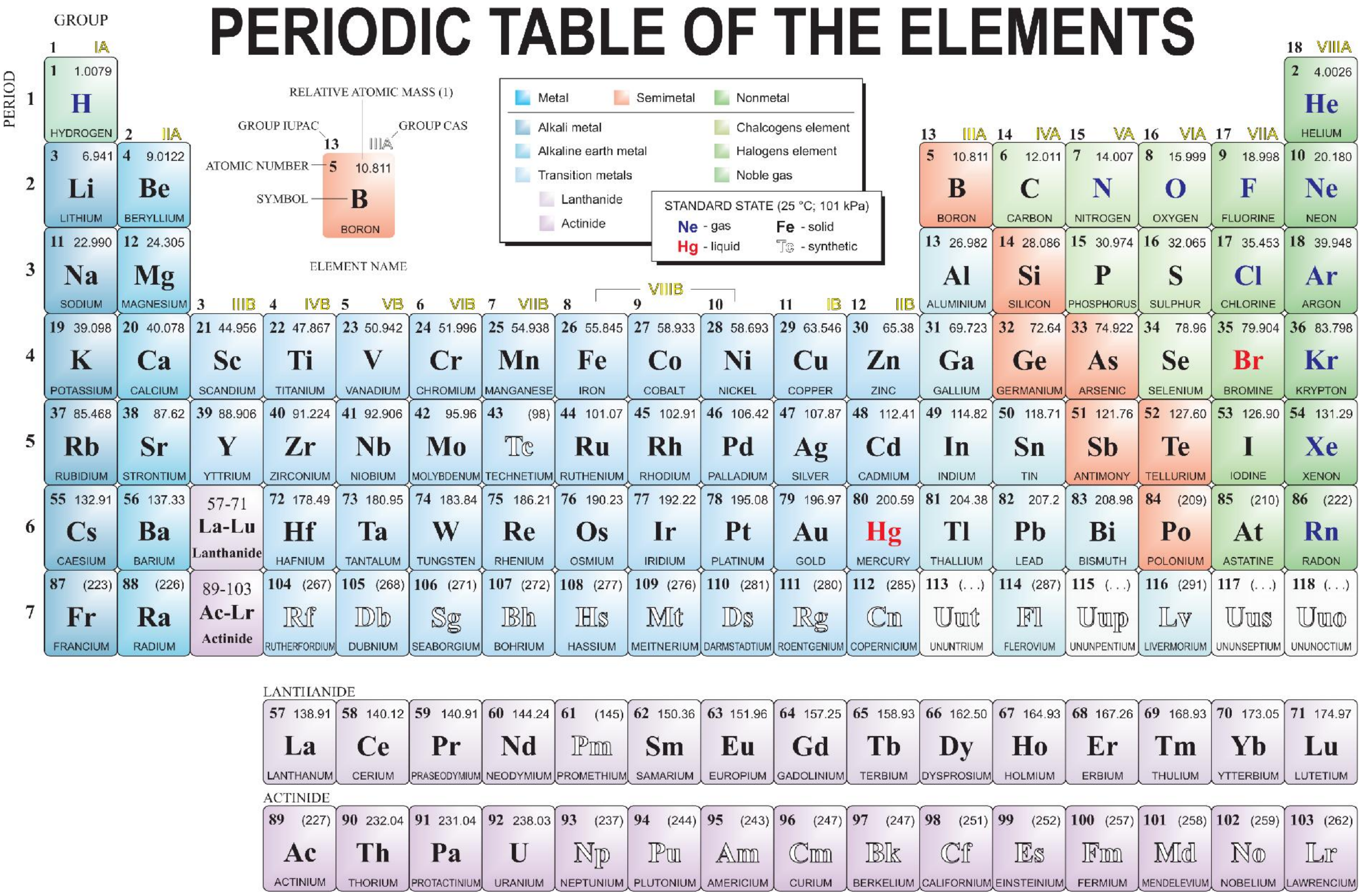
Mendeleev arranged elements in horizontal rows and vertical columns of a table in order of their increasing atomic weights in such a way that the elements with similar properties occupied the same vertical column or group.

2.5 Modern Periodic Law

In 1913, the English physicist, Henry Moseley observed regularities in the characteristic X-ray spectra of the elements. A plot of $\sqrt{v}$ (where v is frequency of X-rays emitted) against **atomic number** (Z) gave a straight line and not the plot of $\sqrt{v}$ vs atomic mass.

Mendeleev's Periodic Law was, therefore, accordingly modified. This is known as the **Modern Periodic Law** and can be stated as :

The physical and chemical properties of the elements are periodic functions of their atomic numbers.



Classification of Modern Periodic Table

s-block
Left
IA, IIA
IA → alkali (ns^1)
IIA → alkaline earth metal (ns^2)
ns^{1-2}

p-block
Right
IIIA-VIIA, VIII
13-18
VIA/16 → chalcogens
VIIA/17 → Halogens
0/18 → Inert gases
ns^2np^{1-6}

d-block
middle
3 → 1 &
IIIB → VIIB
VIII, IB, IIB
transition element
$(n-1)d^{1-10}\ ns^{0-2}$

f-block
Bottom
Inner transition element
$(n-2)f^{1-14}\ (n-1)d^{0-2}\ ns^2$

Nomenclature of elements with Atomic Numbers > 100

The naming of the new elements had been traditionally the privilege of the discoverer and the suggested name was ratified by the IUPAC.

Table : Notation for IUPAC Nomenclature of Elements

Digit	Name	Abbreviation
0	nil	n
1	un	u
2	bi	b
3	tri	t
4	quad	q
5	pent	p
6	hex	h
7	sept	s
8	oct	o
9	enn	e

Table : Nomenclature of Elements with Atomic Number Above 100

Atomic Number	Name	Symbol
101	Unnilunium	Unu
102	Unnibium	Unb
103	Unniltrium	Unt
104	Unnilquadium	Unq
105	Unilpentium	Unp
106	Unnilhexium	Unh
107	Unnilseptium	Uns
108	Unniloctium	Uno
109	Unnilennium	Une
110	Ununnillium	Uun

3. PREDICTION OF BLOCK, PERIOD & GROUP

1. What electronic configuration
2. Block - last e^- enters into which orbital
3. Period - Max value of principal quantum number

   Group - s block - no. of valence electron

   p block - 10 + no. of valence electron

   d block – ns + no. of (n – 1) d e^-

   f block – III B

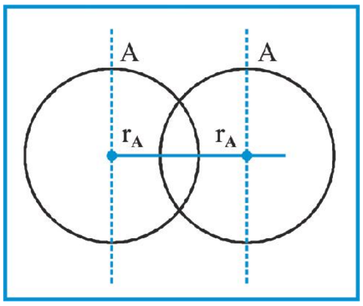
4. PROPERTIES OF AN ELEMENT

4.1 Atomic Radius

We cannot measure the exact size of an isolated atom because its outermost electron have a remote chance of being found quite far from the nucleus. So different types of atomic radius can be used based on the environment of atoms i.e; **covalent** radius, **van der Waals'** radius, **metallic** radius.

4.1.1 Covalent Radius

The half of the distance between the nuclei of two identical atoms joined by single covalent bond in a molecule is known as **covalent radius.**



So covalent radius for A-A

$$\boxed{r_A = \frac{d_{A-A}}{2}}$$

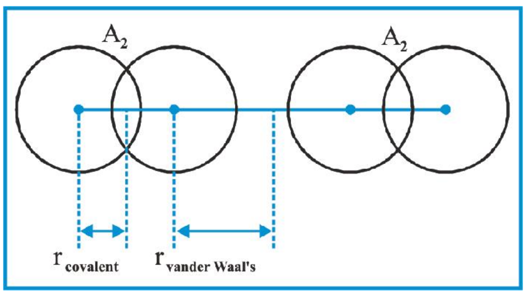
If covalent bond is formed between two different elements then

$$\boxed{d_{A-B} = r_A + r_B - 0.09\,(\chi_A - \chi_B)}$$

where χ_A and χ_B are electronegative of A and B

4.1.2 Vander Waal's Radius

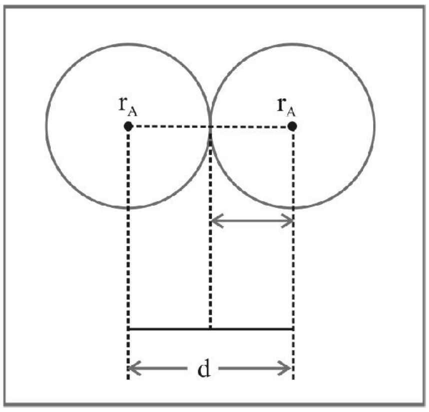
It is half of the internuclear distance between adjacent atoms of the two neighbouring molecules in the solid state.



$$r_{vander\ waal} = \frac{d_{A-A}}{2}$$

4.1.3 Metallic Radius (Crystal radius)

It is one-half of the distance between the nuclei of two adjacent metal atoms in the metallic crystal lattice.



So metallic radius for A-A

$d = r_A + r_A$

$$\boxed{r_A = \frac{d}{2}}$$

$$\boxed{* \ r_{covalent} < r_{metallic} < r_{vander\ waals}}$$

4.2 Variation of Atomic Radii in the Periodic Table

(a) Variation along a period

In general, the covalent and van der Waals radii decrease with increase in atomic number as we move from left to right in a period.

4.3 Atomic Radii

(a) Variation along a period

It is because with in the period the outer electrons are in the same valence shell & the **effective nuclear charge** increases as the atomic number increases resulting in the increased attraction of electrons to the nucleus.

(b) Variation along a group

Atomic radius in a group increase as the atomic number increases. It is because with in the group, the principal quantum number (n) increases and the valence electrons are farther from the nucleus.

(c) Ionic Radius

The removal of an electron from an atom results in the formation of a **cation**, whereas gain of an electron leads to an **anion**.

In general, the ionic radii of elements exhibit the same trend as the atomic radii. A cation is **smaller** than its parent atom because it has fewer electrons while its nuclear charge remains the same. The size of an anion will be **larger** than that of the parent atom because the addition of one or more electrons would result in increased repulsion among the electrons and a decrease in effective nuclear charge. For example, the ionic radius of fluoride ion (F^-) is 136 pm whereas the atomic radius of fluorine is only 64 pm. On the other hand ,the atomic radius of sodium is 186 pm compared to the ionic radius of 95 pm for Na^+.

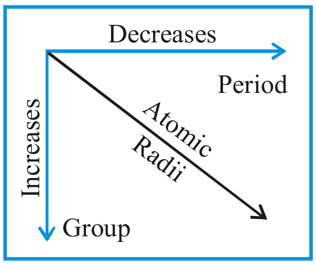
(d) Isoelectronic Species

Isoelectronic species are those which have same number of electrons. For example, O^{2-}, F^-, Na^+ and Mg^{2+} have the same number of electrons (10). Their radii would be different because of their different nuclear charges. The cation with the greater positive charge will have a smaller radius because of the greater attraction of the electrons to the nucleus. Anion with the greater negative charge will have the larger radius. In this case, the net repulsion of the electrons will outweigh the nuclear charge and the ion will expand in size.

Order of atomic radii is

$$\boxed{Mg^{2+} < Na^+ < F^- < O^{2-}}$$

General Trend :



4.4 Ionization Energy

The minimum amount of energy required to remove the electron from the outermost orbit of an isolated atom in the gaseous state is known as ionization energy.

$$M \xrightarrow[-e^-]{IE_1 \text{ (First Ionization Energy)}} M^+ \xrightarrow[-e^-]{IE_2} M^{2+} \xrightarrow[-e^-]{IE_3}$$

$$M^{3+} \xrightarrow[-e^-]{IE_4} M^{4+}$$

IE_1, IE_2, IE_3 and IE_4 are successive ionization energies.

$$\boxed{\begin{aligned} &IE_4 > IE_3 > IE_2 > IE_1 \\ \text{or} \quad &\Delta_i H_4 > \Delta_i H_3 > \Delta_i H_2 > \Delta_i H_1 \end{aligned}}$$

Variation of Ionisation Energy in Periodic Table

(a) Variation along a period

In a period, the value of ionisation enthalpy increases from left to right with breaks where the atoms have somewhat stable configurations. The observed trends can be easily explained on the basis of increased nuclear charge and decrease in atomic radii. Both the factors increase the force of attraction towards nucleus and consequently, more and more energy is required to remove the electrons and hence, ionisation enthalpies increase.

(b) Variation along a group

On moving the group, the atomic size increases gradually due to an addition of one new principal energy shell at each succeeding element. On account of this, the force of attraction towards the valence electrons decreases and hence the ionisation enthalpy value decreases.

4.5 Units of I.E./I.P.

It is measured in units of electron volts (eV) per atom or kilo calories per mole (kcal mol^{-1}) or kilo Joules per mole (kJ mol^{-1}). One electron volt is the energy acquired by an electron while moving under a potential difference of one volt.

1 electron volt (eV) per atom

$= 3.83 \times 10^{-20}$ cal per atom

$= 1.602 \times 10^{-19}$ J per atom (1 cal = 4.184 J)

$= 3.83 \times 10^{-20} \times 6.023 \times 10^{23}$ cal mol^{-1}

$= 23.06$ kcal mol^{-1}

$= 1.602 \times 10^{-19} \times 6.023 \times 10^{23}$ J mol^{-1}

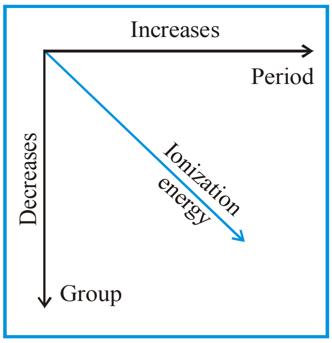
$= 96.49$ kJ mol^{-1}

$\therefore$ 1 electron volt (eV) per atom

$= 23.06$ kcal mol^{-1} $= 96.49$ kJ mol^{-1}

Important Points

* Ionization energy increases with decreasing the size of an atom or an ion
* Ionization energy increases with decreasing screening effect.
* Ionization energy increases with increasing nuclear charge
* Ionization energy increases if atom having half filled and fully filled orbitals
* The penetrating power of orbitals is in the order $s > p > d > f$



Electron Gain Enthalpy

When an electron is added to a neutral gaseous atom (X) to convert it into a negative ion, the enthalpy change accompanying the process is defined as the **Electron Gain Enthalpy ($\Delta_{eg}H$)**. Electron gain enthalpy provides a measure of the ease with which an atom adds an electron to form anion as represented by

$$X(g) + e^- \rightarrow X^-(g)$$

Depending on the element, the process of adding an electron to the atom can be either **endothermic** or **exothermic**. For many elements energy is released when an electron is added to the atom and the electron gain enthalpy is negative. For example, group 17 elements (the **halogens**) have very **high negative electron gain enthalpies** because they can attain stable noble gas electronic configurations by picking up an electron. On the other hand, **noble gases** have large **positive** electron gain enthalpies because the electron has to enter the next higher principal quantum level leading to a very unstable electronic configuration.

Variation of Electron Gain Enthalpy

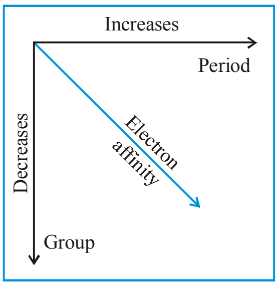
(a) Variation along a period

Electron gain enthalpy becomes more and more negative from left to right in a period. This is due to decrease in size and increase in nuclear charge as the atomic number increases in a period. Both these factors favour the addition of an extra electron due to higher force of attraction by the nucleus for the incoming electron.

(b) Variation along a group

The electron gain enthalpies, in general, become less negative in going down from top to bottom in a group. This is due to increase in size on moving down a group.

This factor is predominant in comparison to other factor, i.e., increase in nuclear charge.



4.6 Electronegativity

The tendency of an atom to attract the shared pair of electrons towards itself is known as its **electronegativity**.

According to **Pauling**, the electronegativity of F is 4.0 and electronegativity of other elements can be calculated as

$(\chi_A - \chi_B) = 0.208\ [E_{A-B} - (E_{A-A} \times E_{B-B})^{1/2}]^{1/2}$

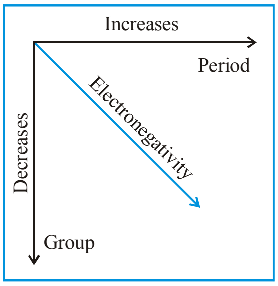
According to **Mulliken**

$$\boxed{\text{Electronegativity} = \frac{IP + EA}{2}}$$

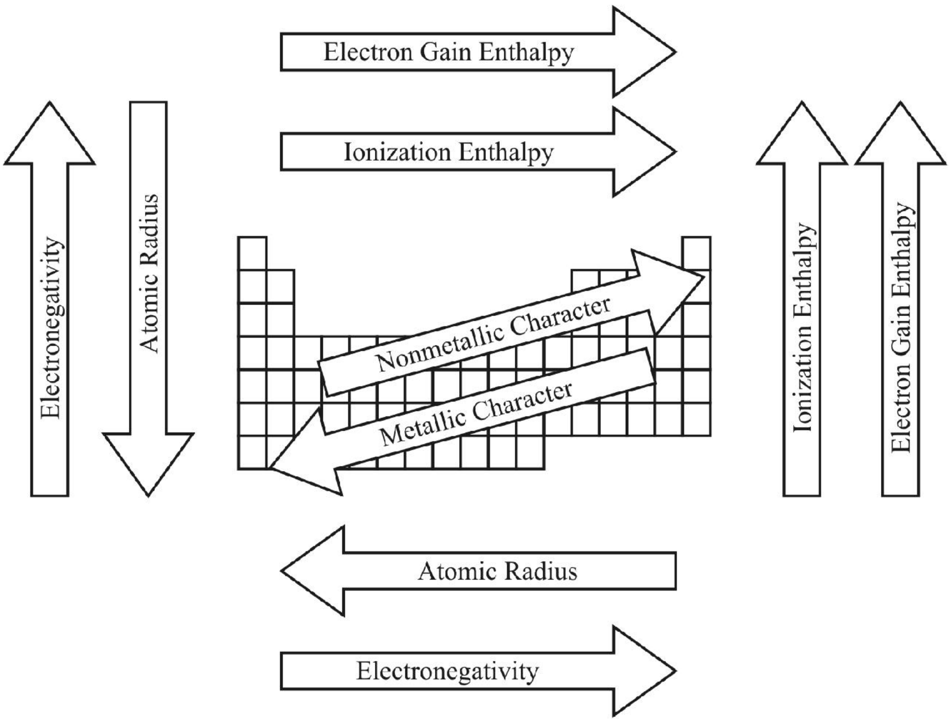
(where IP = Ionization potential, EA = Electron affinity)

If IP and EA are taken in electron volt

* Percentage ionic character = $16\,(\chi_A - \chi_B) + 3.5\,(\chi_A - \chi_B)^2$ where χ_A and χ_B are electronegativities of A and B.
* If the difference in the electronegatives of combining atoms is 1.7, the bond is 50% covalent and 50% ionic.
* If the difference in electronegativities of oxygen and element is very high the oxide shows a basic character.



The periodic trends of elements in the periodic table



4.7 Periodic Trends in Chemical Properties

4.7.1 Periodicity of Valence or Oxidation States

The electrons present in the outermost shell of an atom are called ***valence electrons*** *and the number of these electrons determine the* ***valence*** *or the* ***valency*** *of the atom. It is because of this reason that the outermost shell is also called the* ***valence shell*** *of the atom and the orbitals present in the valence shell are called* ***valence orbitals****.*

In case of representative elements, the valence of an atom is generally *equal to either the number of valence electrons* (s- and p-block elements) or equal to eight minus the number of valence electrons.

Group	1	2	13	14	15	16	17	18
Number of valence electrons	1	2	3	4	5	6	7	8
Valence	1	2	3	4	3, 5	2, 6	1, 7	0, 8

In contrast, transition and inner transition elements, exhibit variable valence due to involvement of not only the valence electrons but d- or f-electrons as well. However, their most common valence are 2 and 3.

Let us now discuss periodicity of valence along a period and within a group.

(a) Variation along a period

As we move across a period from left to right, the number of valence electrons increases from 1 to 8. But the valence of elements, w.r.t. H or O first increases from 1 to 4 and then decreases to zero.

In the formation of Na_2O molecule, oxygen being more electronegative accepts two electrons, one from each of the two sodium atoms and thus shows an oxidation state of –2. On the other hand, sodium with valence shell electronic configuration as $3s^1$ loses one electron to oxygen and is given an oxidation state of +1. Thus, the *oxidation state of an element in a given compound may be defined as the charge acquired by its atom on the basis of electronegativity of the other atoms in the molecule.*

(b) Variation within a group

When we move down the gorup, the number of valence electrons remains the same, therefore, all the elements in a group exhibit the same valence. For example, all the elements of group 1 (alkali metals) have valence one while all the elements of group 2 (alkaline earth metals) exhibit a valence of two.

Noble gases present in group 18 are zerovalent, i.e., their valence is zero since these elements are **chemically inert.**

4.7.2 Anomalous Properties of Second Period Elements

It has been observed that *some elements of the second period show similarities with the elements of the third period placed diagonally to each other, though belonging to different groups*. For example, lithium (of group 1) resembles magnesium (of group 2) and beryllium (of group 2) resembles aluminium (of group 13) and so as. *This similarity in properties of elements placed diagonally to each other is called* **diagonal relationship.**

$$\begin{tikzcd} \textbf{Group 1} & \textbf{Group 2} & \textbf{Group 13} & \textbf{Group 14} \\ \mathrm{Li} \arrow[dr] & \mathrm{Be} \arrow[dr] & \mathrm{B} \arrow[dr] & \mathrm{C} \\ \mathrm{Na} & \mathrm{Mg} & \mathrm{Al} & \mathrm{Si} \end{tikzcd}$$

The anomalous behaviour is due to their small size, large charge/radius ratio and high electronegativity of the elements. In addition, the first member of group has only four valence orbitals (2s and 2p) available for bonding, whereas the second member of the groups have nine valence orbitals (3s, 3p, 3d). As a consequence of this, the maximum covalency of the first member of each group is 4 (e.g., boron can only form $[BF_4])^-$, whereas the other members of the groups can expand their valence shell to accommodate more than four pairs of electrons e.g., aluminium forms $[AlF_6]^{3-}$. Furthermore, the first member of p-block elements displays greater ability to form pπ-pπ multiple bonds to itself (e.g., C=C, C≡C, N=N, N≡N) and to other second period elements (e.g., C=O, C=N, C≡N, N=O) compared to subsequent members of the same groups.

4.7.3 Periodic Trends and Chemical Reactivity

Reactivity of Metals

The reactivity of metals is measured in terms of their tendency to lose electrons from their outermost shell.

In a period

The tendency of an element to lose electrons decreases in going from left to right in a period. So, the reactivity of metals decreases in a period from left to right. For example, the reactivity of third period elements follows the order.

$$\underset{\textbf{more reactive}}{\mathrm{Na}} > \mathrm{Mg} > \underset{\textbf{reactive}}{\mathrm{Al}}$$

In a group

The tendency to lose electrons increases as we go down a group. So, the reactivity of metals increases down the group. Thus, in group 1, the reactivity follows the order.

$$\underset{\textbf{Least reactive}}{\mathrm{Li}} < \mathrm{Na} < \mathrm{K} < \mathrm{Rb} < \underset{\textbf{Most reactive}}{\mathrm{Cs}}$$

—— Reactivity increases ⟶

Reactivity of Non-Metals

The reactivity of a non-metal is measured in terms of its tendency to gain electrons to form an anion.

In a period

The reactivity of non-metals increases from left to right in a period. During reaction, non-metals tend to form anions. For example, in the second period, the reactivity of non-metals increases in the order.

$$\underset{\text{Less reactive}}{C} < N < O < \underset{\text{More reactive}}{F}$$

—— Reactivity increases ⟶

In a group

The reactivity of non-metals in a group decreases as we go down the group. This is because the tendency to accept electrons decreases down the group. The reactivity of halogens follows the order

$$\underset{\text{Most reactive}}{F} > Cl > Br > \underset{\text{Least reactive}}{I}$$

—— Reactivity decreases ⟶

The normal oxide formed by the element on extreme left is the **most basic** (e.g., Na_2O) whereas that formed by the element on extreme right is the **most acidic** (e.g., Cl_2O_7). Oxides of elements in the centre are **amphoteric** (e.g., Al_2O_3, As_2O_3) or **neutral** (e.g., CO, NO, N_2O). Amphoteric oxides behave as acidic with bases and as basic with acids, whereas neutral oxides have no acidic or basic properties.

4.7.4 Inert Pair Effect

In groups 13-16, as we move down the group, the tendency of s-electrons of the valence shell to participate in bond formation decreases. This means that lower oxidation state becomes more stable.

Reason: As we go down these groups, the increased nuclear charge outweighs the effect of the corresponding increase in atomic size. The s-electrons thus become more tightly held (more penetrating) and hence more reluctant to participate in bond formation. Hence, the lower oxidation state becomes more stable.

5. SUMMARY AND IMPORTANT POINTS TO REMEMBER

1. Mendeleev's periodic table was based on atomic masses of the elements. When Mendeleev presented the periodic table, only 63 elements were known. He left 29 places in the table for unknown elements.

2. Modern Mendeleev periodic table is based on atomic numbers of the elements. The modern periodic law is : **"The physical and chemical properties of the elements are periodic function of their atomic numbers".**

   The horizontal row in the periodic table is called a **period** and vertical column is called **group**. There are seven periods and nine groups in the modern Mendeleev periodic table.

3. The long or extended form of periodic table consists of seven periods and eighteen vertical columns (groups or families). The elements in a period have same number of energy shells, i.e., principal quantum number (n). These are numbered 1 to 7.

1st period	1s	2 elements
2nd period	2s 2p	8 elements
3rd period	3s 3p	8 elements
4th period	4s 3d 4p	18 elements
5th period	5s 4d 5p	18 elements
6th period	6s 4f 5d 6p	32 elements
7th period	7s 7f 6d 7p	32 elements
	Total	*118 elements

   At present 114 elements are known.

   In a vertical column (group), the elements have similar valence shell electronic configuration and therefore exhibit similar chemical properties.

4. There are four blocks of elements: s-, p-, d- and f-block depending on the orbital which gets the last electron. The general electronic configuration of these blocks are :

   s-block : [Noble gas] $ns^{1 \text{ or } 2}$. However, hydrogen has 1s1 configuration.

   p-block : [Noble gas] ns^2np^{1-6}

   d-block : [Noble gas] $(n-1)d^{1-10}ns^{1 \text{ or } 2}$

   f-block : [Noble gas] $(n-2)f^{1-14}(n-1)d^{0 \text{ or } 1}ns^2$

   s-block elements occupy IA(1) and IIA(2) groups, i.e., extreme left portion of the periodic table.

   p-block elements occupy IIIA(13), IVA(14), VA(15), VIA(16), VIIA(17) and VIIIA(18) groups, i.e., right portion of the periodic table.

   d-block elements occupy IIIB(3), IVB(4), VB(5), VIB(6), VIIB(7), VIIB(8, 9 and 10), IB(11) and IIB(12) groups, i.e., central portion of the periodic table. There are four d-block series, i.e., 3d series, 4d series, 5d series and 6d series, each consisting of ten elements, i.e., in all forty d-block elements are present in periodic table.

   f-block elements are accommodated in two horizontal rows below the main periodic table, each row consists of 14 elements, i.e., 28 f-block elements are present in periodic table. The elements in first row are termed 4f-elements or rare earth or lanthanides while the elements of second row are termed 5f-elements or actinides.

5. The elements are broadly divided into three types :

(i) **Metals** comprise more than 78% of the known elements. s-block, d-block and f-block elements are metals. The higher members of p-block are also metals.

(ii) **Non-metals** are less than twenty. (C, N, P, O, S, Se, H, F, Cl, Br, I, He, Ne, Ar, Kr, Xe and Rn are non-metals).

(iii) Elements which lie in the border line between metals and non-metals are called **semimetals** or **metalloids**. B, Si, Ge, As, Sb, Te, Po and At are regarded metalloids.

6. IUPAC given a new scheme for assigning a temporary name to the newly discovered elements. The name is derived directly from the atomic number of the elements. However, IUPAC has accepted the following names of the elements from atomci numbers 104 to 110.

Rutherfordium (Rf), 104	Dubnium (Db), 105	Seaborgium (Sg) 106
Bohrium (Bh), 107	Hassium (Hs), 108	Meitnerium (Mt), 109
Darmstadtium (Ds) 110		

   The temporary names of the elements discovered recently are :

Ununumium (Uuu), 111	Ununbium (Uub) 112	
Ununquadium (Uuq) 114	and	Ununhexium (Uuh) 116

7. The recurrence of similar properties of the elements after certain definite intervals when the elements are arranged in order of increasing atomic numbers in the periodic table is termed **periodicity**. The cause of periodicity is the repetition of similar electronic configuration of the atom in the valence shell after certain definite intervals. These definite intervals are 2, 8, 8, 18, 18 and 32. These are known as magic number.

   Periodicity is observed in a number of properties which are directly or indirectly linked with electronic configuration.

(i) Effective nuclear charge increases across each period.

(ii) Atomic radii generally decrease across the periods.

(iii) Atomic radii generally increase on moving from top to bottom in the groups.

(iv) Atomic radius is of three types :

(a) **Covalent radius :** It is half of the distance between the centres of the nuclei of two similar atoms joined by a single covalent bond. This is generally used for non-metals.

(b) **Crystal or metallic radius :** It is half of the internuclear distance between two nearest atoms in the metallic lattice. It is generally used for metals.

(c) **van der Waals' radius :** It is half of the internuclear distance between the nearest atoms belonging to two adjacent molecules in solid state.

van der Waals' radius > Metallic radius > Covalent radius (for an atom)

(v) Cations are generally smaller than anions.

(vi) Cations are smaller and anions are larger than neutral atoms of the elements.

Cation size < Neutral atom size < Anion size

(vii) Elements of 2nd and 3rd transition series belonging to same vertical columns are similar in size and properties due to **lanthanide contraction**.

(viii) The first element is each group of the representative elements shows abnormal properties, i.e., differs from other elements of the group because of much smaller size of the atom.

(ix) The ions having same number of electrons but different nuclear charge are called **isoelectronic ions**.

Examples,

(a) N^{3-}, O^{2-}, F^{-}, Na^{+}, Mg^{2+}, Al^{3+}

(b) P^{3-}, S^{2-}, Cl^{-}, K^{+}, Ca^{2+}, Sc^{3+}

In isoelectronic ions, the size decreases if Z/e increases i.e., greater the nuclear charge, smaller is the size of the ion.

(x) The energy required to remove the most loosely held electron from the gaseous isolated atom is termed ionisation enthalpy.

(xi) Ionisation enthalpy values generally increase across the periods.

(xii) Ionisation enthalpy values generally decrease down the group.

(xiii) Removal of electron from filled and half filled shells requires of higher energy. For example, the ionisation enthalpy of nitrogen is higher than oxygen. Be, Mg and noble gases have high values.

(xiv) Metals have low ionisation enthalpy values while non-metals have high ionisation enthalpy values.

(xv) Successive ionisation enthalpies of an atom have higher values.

$IE_{I} < IE_{II} < IE_{III} \ldots$

(xvi) The enthalpy change taking place when an electron is added to an isolated gaseous atom of the element is called electron gain enthalpy. The first electron gain enthalpy of most of the elements is negative as energy is released in the process but the values are positive or near zero in case of the atoms having stable configuration such as Be, Mg, N, noble gases, etc.

(xvii) Electron gain enthalpy becomes more negative from left to right in a period and less negative from top to bottom in a group.

(xviii) Successive electron gain enthalpies are always positive.

(xix) The elements with higher ionisation enthalpy have higher negative electron gain enthalpy.

(xx) Electronegativity is the tendency of an atom to attract the shared pair of electrons towards itself in a bond.

(xxi) Electronegativity increases across the periods and decreases down the groups.

(xxii) Metals have low electronegativities and non-metals have high electronegativities.

(xxiii) Metallic character decreases across the periods and increases down the group.

(xxiv) Valence of an element belonging to s- and p- block (except noble gases) is either equal to the number of valence electrons or eight minus number of valence electrons.

(xxv) The **reducing nature** of the elements decreases across the period while **oxidising nature** increases.

(xxvi) The **basic character** of the oxides decreases while the **acidic character** increases in moving from left to right in a period.

6. SOME IMPORTANT FACTS ABOUT ELEMENTS

(i) Bromine is a non-metal which is liquid at room temperature.

(ii) Mercury is the only metal that is liquid at room temperature.

(iii) Gallium (m.pt. 29.8ºC), caesium (m.pt. 28.5ºC) and francium (m.pt. 27ºC) are metals having low melting points.

(iv) Tungsten (W) has the highest melting point (3380ºC) among metals.

(v) Carbon has the highest melting point (4100ºC) among non-metals.

(vi) Oxygen is the most abundant element on the earth.

(vii) Aluminium is the most abundant metal.

(viii) Iron is the most abundant transition metal.

(ix) Highest density is shown by osmium (22.57 g cm^{-3}) or iridium (22.61 g cm^{-3}).

(x) Lithium is the lightest metal. Its density is 0.54 g cm^{-3}.

(xi) Silver is the best conductor of electricity.

(xii) Diamond (carbon) is the hardest natural substance.

(xiii) Francium has the highest atomic volume.

(xiv) Boron has the lowest atomic volume.

(xv) The most abundant gas in atmosphere is nitrogen.

(xvi) Fluorine is the most electronegative element.

(xvii) Chlorine has the maximum negative electron gain enthalpy.

(xviii) Helium has the maximum ionisation enthalpy.

(xix) Cesium or francium has the lowest ionisation enthalpy.

(xx) Helium and francium are smallest and largest atoms respectively.

(xxi) H^- and I^- ions are the smallest and largest anions respectively.

(xxii) H^+ and Cs^+ ions are the smallest and largest cations respectively.

(xxiii) Cesium is the most electropositive element.

(xxiv) Element kept in water is phosphorus, P_4 (white or yellow).

(xxv) Element kept in kerosene are Na, K, Rb, Cs, etc.

(xxvi) Iodine is the element which sublimes.

(xxvii) Hydrogen is the most abundant element in the universe.

(xxviii) Only ozone is the coloured gas with garlic smell.

(xxix) Metalloids have electronegativity values closer to 2.0.

(xxx) First synthetic (i.e., man-made) element is technetium (At. No. 43).

(xxxi) Most poisonous metal-Plutonium.

(xxxii) Rarest element in earth's crust-Astatine.

(xxxiii) The elements coming after uranium are called transuranic elements. The elements with Z = 104 – 112, 114 and 116 are called trans-actinides or super heavy elements. All these elements are synthetic, i.e., man-made elements. These are radioactive elements and not found in nature.

(xxxiv) The elements ruthenium (Ru), germanium (Ge), polonium (Po) and americium (Am) were named in honour of the countries named **Ruthenia (Russia), Germany, Poland** and **America,** respectively.

(xxxv) The members of the actinide series are radioactive and majority of them are not found in nature.

(xxxvi) The element rutherfordium (Rf, 104) is also called Kurchatovium (Ku) and element dubnium (Db, 105), is also called hahnium.

(xxxvii) Promethium (Pm, 61) a member of lanthanide series is not found in nature. It is a synthetic element.

(xxxviii) Special names are given to the members of these groups in periodic table.

Group 1	or	IA	Alkali metals
Group 2	or	IIA	Alkaline earth metals
Group 15	or	VA	Pnicogens
Group 16	or	VIA	Chalcogens
Group 17	or	VIIA	Halogens
Group 18	or	VIIIA (zero)	Inert or noble gases

SOLVED EXAMPLES

Example – 1

What is the basic difference in approach between the Mendeleev's periodic law and the modern periodic law ?

Sol. According to Mendeleev, the properties of the elements are a periodic function of their atomic weights, while according to modern periodic law, the properties of the elements are periodic functions of their atomic numbers.

Example – 2

On the basis of quantum numbers, justify that the sixth period of the periodic table should have 32 elements.

Sol. Sixth period corresponds to n = 6. In this period 16 orbitals, viz. one 6s, seven 4f, five 5d and three 6p orbitals are filled. These sixteen orbitals can accommodate 32 elements. So, there are 32 elements in the sixth period.

Example – 3

What do you understand by isoelectronic species ? Name the species that will be isoelectronic with each of the following atoms or ions.

(i) F^- **(ii) Ar**

(iii) Mg^{2+} **(iv) Rb^+**

Sol. Ions of different elements which have the same number of electrons but different magnitude of the nuclear charge are called isoelectronic ions.

(i) F^- has 10 (9 + 1) electrons. Therefore, the species nitride ion, N^{3-} (7 + 3) ; oxide ion ; O^{2-} (8 + 2), neon, Ne (10 + 0) ; sodium ion, Na^+ (11 – 1) ; magnesium ion, Mg^{2+} (12 – 2) ; aluminium ion, Al^{3+} (13 – 3) etc. each one of which contains 10 electrons, are isoelectronic with it.

(ii) Ar has 18 electrons. Therefore, the species phosphide ion, P^{3-} (15 + 3), sulphide ion ; S^2 (16 + 2) ; chloride ion, Cl^- (17 + 1), potassium ion, K^+ (19 – 1), calcium ion, Ca^{2+} (20 – 2), etc. each one of which contains 18 electrons, are isoelectronic with it.

(iii) Mg^{2+} has 10 (12 – 2) electrons, therefore, the species N^{3-}, O^{2-}, F^-, Ne, Na^+, Al^{3+}, etc. each one of which contains 10 electrons, are isoelectronic with it.

(iv) Rb^+ has 36 (37 – 1) electrons. Therefore, the species bromide ion, Br^- (35 + 1), krypton, Kr (36 + 0) and strontium Sr^{2+} (38 – 2) each one of which has 36 electrons, are isoelectronic with it.

Example – 4

Give four examples of species which are isoelectronic with Ca^{2+}.

Sol. K^+, Cl^-, S^{2-} or P^{3-} are isoelectronic with Ca^{2+}.

Example – 5

What is periodicity ? What is the cause of periodicity ?

Sol. When elements are arranged in the increasing order of atomic number, elements having similar properties recur at regular intervals in the periodic table. This type of property is called periodicity.

The cause of the periodicity in properties is the same outermost electronic configuration coming at regular intervals.

Example – 6

(a) What is modern periodic law ? Discuss the main features of long form of periodic table.

(b) Give the general electronic configuration of s, p, d & f-block elements.

Sol. (a) The physical and chemical properties of the elements are periodic functions of their atomic numbers. The main features of long form of periodic table are as follows :

1. The aufbau (build up) principle and the electronic configuration of atoms provide a theoretical foundation for the periodic classification.
2. The long form of the periodic table consists of horizontal rows called periods and vertical columns called groups.
3. There are altogether seven periods. The period number corresponds to the highest principal quantum number (n) of the elements in the period.
4. The first period contains 2 elements. The subsequent periods consists of 8, 8, 18, 18 and 32 elements, respectively. The seventh period is incomplete and like the sixth period would have a theoretical maximum (on the basis of quantum numbers) of 32 elements.
5. In this form of the Periodic Table, 14 elements of both sixth and seventh periods (lanthanoids and actinoids, respectively) are placed in separate panels at the bottom.
6. Elements having similar outer electronic configurations in their atoms are arranged in vertical columns referred to as groups or families. There are in all 18 vertical column or groups.
7. The elements of groups 1 (alkali metals), 2 (alkaline earth metals) and 13 to 17 are called the main group elements. These are also called typical or representative or normal elements.
8. The elements of group 3 to 12 are called transiation elements.
9. lanthanoids & actinoids are together referred to as inner transition elements.

(b) (i) General outer electronic configuration of s-block elements is $\mathbf{ns^{1-2}}$ i.e., either ns^1 or ns^2.

(ii) General outer electronic configuration of p-block elements is $\mathbf{ns^2np^{1-6}}$.

(iii) General outer electronic configuration of d-block elements is $\mathbf{(n-1)\ d^{1-10}\ ns^{0-2}}$.

(iv) General outer electronic configuration of f-block elements is $\mathbf{(n-2)\ f^{1-14}\ (n-1)\ d^{0-1}\ ns^2}$.

Example – 7

Give general electronic configuration of f-block elements and explain characteristic properties of elements in lanthanide and actinide series.

Sol. (i) General electronic configuration of f-block elements is $\mathbf{(n\text{-}2)\ f^{1\text{-}14}\ (n\text{-}1)\ d^{0\text{-}1}\ ns^2}$.

(ii) The last electron in these elements enter f-orbital of the prepenultimate shell of an atom.

(iii) It includes the lanthanide and actinide series of group 3 (IIIB).

(iv) These elements are placed separatey in two rows, at the bottom of the periodic table.

Characteristic properties of elements in the lanthanide and actinide series are :

(i) Elements from the series show very little difference in their chemical reactivity.

(ii) Elements are metallic, electropositive with high melting point and boiling point.

(iii) Most of their compounds are coloured in solid state as well as in aqueous solution.

(iv) The ability to form complex compounds is comparatively less.

(v) The elements of both the series are paramagnetic as they have unpaired electrons.

(vi) As atomic number increases, atomic size slightly decreases across individual series. The decrease in atomic size is called lanthanide contraction for lanthanide series and actinide contraction for actinide series.

(vii) The common oxidation state of lanthanide and actinide series is +3.

(viii) They also show variable oxidation state of +2, +4, +5 and +6. Hence exhibit good catalytic activity.

(ix) Some of the elements are **radioactive**. (Elements heavier than Uranium, do not occur in nature but are obtained artificially through nuclear reactions. These elements are called **transuranic elements**.)

Example – 8

Define atomic radius. Explain various factors affecting it ?

Sol. Atomic radius is defined as the distance of valence shell of electrons from the centre of the nucleus of an atom.

Factors affecting atomic radius :

(i) **No. of shells :** The atomic radius increases with the increase in the no. of the shells.

atomic radius α no of shells

(ii) **Nuclear charge :** Atomic radius decreases with the increase in the Nuclear charge. Due to high nuclear charge, the nucleus attracts the electrons towards itself thereby reducing its own size

$$\boxed{\text{atomic radius } \alpha \frac{1}{\text{Nuclear charge}}}$$

(iii) **Shielding or screening effect :** Atomic radius increases with the increase in the shielding effect. This is because the electrons presents between the Nucleus and the valence shell shields the valence electrons from the Nucleus i.e. it reduces the force of attraction between the Nucleus and the Vaence electorns.

$$\boxed{\text{atomic radius } \alpha \text{ shielding effect}}$$

Example – 9

Explain why cations are smaller and anions are larger in radii than their parent atoms ?

Sol. The ionic radius of a cation is always smaller than the parent atom because the loss of one or more electrons increases the effective nuclear charge. As a result, the force of attraction of nucleus for the electrons increases and hence the ionic radii decrease. In contrast, the ionic radius of an anion is always larger than its parent atom because the addition of one or more electrons decreases the effective nuclear charge. As a result, the force of attraction of the nucleus for the electrons decreases and hence the ionic radii increase.

Example – 10

Consider the following species. N^{3-}, O^{2-}, F^{-}, Na^{+}, Mg^{2+} and Al^{3+}.

(i) What is common in them ?

(ii) Arrange them in order of increasing ionic radii ?

Sol. (a) Each one of these ions contains 10 electrons and hence all are isoelectronic ions.

(b) The ionic radii of isoelectronic ions decrease with the increase in the magnitude of the nuclear charge. For example, consider the isoelectronic ions : N^{3-}, O^{2-}, F^{-}, Na^{+}, Mg^{2+} and Al^{3+}. All these ions have 10 electrons but their nuclear charges increase in the order :

N^{3-} (+ 7), O^{2-} (– 8), F^{-} (+ 9), Na^{+} (+ 11), Mg^{2+} (+ 12) and Al^{3+} (+ 13). Therefore, their ionic radii decrease in the order :

N^{3-} $O^{2-} > F^{-} > Na^{+} > Mg^{2+}$ Al^{3+}.

Example – 11

Arrange following in increasing order of size
Ca^{2+}, S^{2-}, P^{3-}, K^{+}

Sol. Ions of different elements which have the same number of electrons are called isoelectronic ions.

The ionic radii of isoelectronic ions decrease with the increase in the magnitude of the nuclear charge.

All these ions have 18 electrons but nuclear charges increases in the order :

P^{3-} (+ 15), S^{2-} (16), K^{+} (19), Ca^{2+} (20)

$\therefore$ their ionic radii increase in the order

$Ca^{2+} < K^{+} < S^{2-} < P^{3-}$

Example – 12

Out of Cl^{-} & Cl which one is larger & why ?

Sol. Cl^{-} +17))) 2 8 8 Cl +17))) 2 8 7

In chloride ion, addition of one more electron in outermost shell decreases the effective nuclear charge. As a result, the force of attraction of the nucleus for the electrons decreases and hence the ionic radii increase.

$\therefore$ Size of chloride ion is more than chlorine atom.

Example – 13

Among the elements Li, K, Ca, S and Kr, which one is expected to have the lowest first ionization enthalpy and which the highest first ionization enthalpy ?

Sol. K has the lowest first ionization energy. Kr has the highest first ionization energy.

Example – 14

Among the second period elements, the actual ionization energies are in the order :

Li < B < Be < C < O < N < F < Ne.

Explain why (i) Be has higher $\Delta_i H$ than B (ii) O has lower $\Delta_i H$ than N and F ?

Sol. (i) The ionization enthalpy, among other things, depends upon the type of electron to be removed from the same principal shell. In case of Be ($1s^2\,2s^2$) the outermost electron is present in 2s-orbital while in B ($1s^2\,2s^2\,2p^1$) it is present in 2p-orbital. Since 2s-electrons are more strongly attracted by the nucleus than 2p-electrons, therefore, lesser amount of energy is required to knock out a 2p-electron than a 2s-electron. Consequently, $\Delta_i H$ of Be is higher than that $\Delta_i H$ of B.

(ii) The electronic configuration of N ($1s^2\,2s^2\,2p^1_x\,2p^1_y\,2p^1_z$) in which 2p-orbitals are **exactly half-filled** is more stable than the electronic configuration of O ($1s^2\,2s^2\,2p^2_x\,2p^1_y\,2p^1_z$) in which the 2p-orbitals are neither exactly half-filled nor completely filled. Therefore, it is difficult to remove an electron from N than from O. As a result, $\Delta_i H$ of N is higher than that of O. Further, the electronic configuration of F is $1s^2\,2s^2\,2p^2_x\,2p^2_y\,2p^1_z$. Because of higher nuclear charge (+9), the first ionization enthalpy of F is higher than that of O. Further, the effect of increased nuclear charge outweights the effect of stability due to exactly half-filled orbitals, therefore, the $\Delta_i H$ of N and O are lower than that of F.

Example – 15

Would you expect the first ionization enthalpies of two isotopes of the same element to be same or different ? Justify your answer.

Sol. Ionization enthalpy, among other things, depends upon the electronic configuration (number of electrons), and nuclear charge (number of protons). Since the isotopes of an element have the same electronic configuration and same nuclear charge, they are expected to have same ionization enthalpy.

Example – 16

I.P. of Mg is greater than I.P. of Al.

Sol. (i) Ionisation potential may be defined as the amount of energy required to remove the most lossely bound electron from an isolated gaseous atom or ion.

(ii) Electronic configuration of Mg is $1s^2\,2s^2\,2p^6\,3s^2$.It has a completely filled valence orbital (i.e. 3s).

(iii) Whereas electronic configuration of Al is $1s^2\,2s^2\,2p^6\,3s^2\,3p^1$. It has a incompletely filled 3p orbital.

(iv) As completely filled and half filled orbitals are more stable than incompletely filled orbitals, the amount of energy required to remove the valence electron from the incompletely filled 3p orbital of Al is less compared to the completely filled 3s orbital of Mg.

Hence IP of Mg is greater than I.P. of Al.

Example – 17

I.P. of P is greater than I.P. of S.

Sol. (i) Ionisation potential may be defined as the amount of energy required to remove the most loosely bound electron from an isolated gaseous atom or ion.

(ii) Electronic configuration of P is $1s^2\,2s^2\,2p^6\,3s^2\,3p^3$. It has half filled 3p orbital.

(iii) Electronic configuration of S is $1s^2\,2s^2\,2p^6\,3s^2\,3p^4$. It has incompletely filled 3p orbital.

(iv) As completely filled and half filled orbitals are more stable than incompletely filled orbitals, the amount of energy required to remove the valence electron from incompletely filled 3p orbital of S is less compared to the half filled 3p orbital of P.

Hence I.P. of P is greater than I.P. of S.

Example – 18

Third I.P. of Mg is maximum in third row.

Sol. (i) Ionisation potential may be defined as the amount of energy required to remove the most loosely bound electron from an isolated gaseous atom or ion.

Electronic configuration of Mg is $1s^2\,2s^2\,2p^6\,3s^2$

First I.P. involves removal of one electron from $3s^2$

$Mg \rightarrow Mg^{+1} + e^{-1}$

Second I.P. involves removal of one electron from $3s^1$

$Mg^{+1} \rightarrow Mg^{+2} + e^{-1}$

Third I.P. involves removal of one electron from $2p^6$

$Mg^{+2} \rightarrow Mg^{+3} + e^{-1}$

As the third electron is removed form completely filled $2p^6$ orbital, the energy required is maximum.

Hence IP of Mg is maximum in 3[rd] row.

Example – 19

I.P. of noble gases is maximum.

Sol. 'Ionisation energy or Ionisation potential of an element may be defined as the amount of energy required to remove the most loosely bound electron from an isolated gaseous atom.'

The general electronic configuration of noble is ns^2np^6. Since noble gas has **completely filled p-orbital**, the energy required is maximum.

Hence I.P. of noble gases is maximum.

Example – 20

Among the elements B, Al, C and Si

(i) Which element has the highest first ionisation enthalpy ?

(ii) Which element has the most metallic character ?

Justify your answer in each case.

Sol. Arrange the elements B, Al, C and Si into different groups and periods in order of their increasing atomic numbers, we have,

Group →	13	14
Period 2	B	C
Group 3	Al	Si

(i) Since ionization enthalpy increases along a period and decreases down a group, therefore, **C** has the highest first ionization enthalpy.

(ii) Since metallic character increases down a group and decreases along a period, therefore, **Al**, is the most metallic element.

Example – 21

Why is ionization enthalpy of nitrogen greater than that of oxygen ?

Sol. Nitrogen has stable exactly half-filled p-orbitals.

Example – 22

Δ_iH_1 value of Mg is more as compare to that of Na while its Δ_iH_2 value is less. Explain.

or

How would you explain the fact that the first ionization enthalpy of sodium is lower than that of magnesium but its second ionization enthalpy is higher than that of magnesium ?

Sol. The electronic configurations of Na and Mg are

Na : $1s^2\ 2s^2\ 2p^6\ 3s^1$ and Mg : $1s^2\ 2s^2\ 2p^6\ 3s^2$.

Thus, the first electron in both the cases has to be removed from the 3s-orbital but the nuclear charge of Na (+ 11) is lower than that of Mg (+ 12), therefore, the first ionization energy of sodium is lower than that of magnesium. After the loss of first electron, the electronic configuration of Na^+ is $1s^2\ 2s^2\ 2p^6$. Here, the electron is to be removed from inert (neon) gas configuration which is very stable and hence removal of second electron from sodium is very difficult. However, in case of magnesium, after the loss of first electron, the electronic configuration of Mg^+ is $1s^2\ 2s^2\ 2p^6\ 3s^1$. (Here, the electron is to be removed from a 3s orbital which is much easier than to remove an electron from inert gas configuration. Therefore, the second ionization enthalpy of sodium is higher than that magnesium.)

Example – 23

Explain why ionization enthalpies decrease down the group of the periodic table ?

Sol. Ionization enthalpies decrease down the group of the periodic table because inner shell increases. As the distance of the outer electrons from the nucleus increases with increase in atomic radius, the attractive force on the outer electrons decreases & hence lesser amount of energy is required to knock them out.

Example – 24

Why does the first ionisation enthalpy increase as we go from left to right across a given period of the periodic table ?

Sol. The value of ionisation enthalpy increases with the increase in atomic number across the period. This is due to the fact that in moving across the period from left to right :

(i) nuclear charge increases regularly by one unit.

(ii) progressive addition of electrons occurs in the same level.

(iii) atomic size decreases.

This is due to the gradual increase in nuclear charge and with the simultaneous decrease in atomic size the electrons are more and more tightly bound to the nucleus. This results in the gradual increase in ionisation energy across the period.

Example – 25

Would you expect the second electron gain enthalpy of O as positive, more negative or less negative than the first ? Justify your answer.

Sol. The second electron gain enthalpy of O is positive as explained below :

When an electron is added to O atom to form O^- ion, energy is released. Thus, first electron gain enthalpy of O is negative.

$O(g) + e^-(g) \rightarrow O^-(g)$; $\Delta_{eg} H = -141$ kJ mol^{-1}

But when another electron is added to O^- to form O^{2-} ion, energy is absorbed to overcome the strong electrostatic repulsion between the negatively charged O^- ion and the second electron being added. Thus, the second electron gain enthalpy of oxygen is positive.

$O^-(g) + e^-(g) \rightarrow O^{2-}(g)$; $\Delta_{eg} H = +780$ kJ mol^{-1}

Example – 26

Among the elements of the third period of Na to Ar pick out the element

(i) with the highest first ionization energy

(ii) with the largest atomic radius

(iii) that is the most reactive non-metal

(iv) that is the most reactive metal.

Sol. (i) Argon, (ii) Na, (iii) Chlorine, (iv) Sodium.

Example – 27

'Electron affinity of fluorine is less than that of chlorine'. Explain.

Sol. (i) Def : Electron gain enthalpy or electron affinity is defined as the amount of energy released, when neutral gaseous atom, accepts an electron to form an anion.

(ii) The electronic cofiguration of fluorine is $1s^2 2s^2 2p^5$, while that of chlorine, it is $1s^2 2s^2 2p^6 3s^2 3p^5$. In both the elements there are 7 electrons in their outermost shell. The size of F-atom is smaller than Cl-atom.

(iii) In fluorine, 2p-orbitals are **compact** and **closer** to the nucleus. Thus, the screening effect is very low. Hence there is electron-electron repulsion in the valence shell. Thus, when an electron is added to the p-orbital of a fluorine. Thus, when an electron is added to the p-orbital of a fluorine it experiences less attraction and hence less energy is liberated to form fluoride ion.

(iv) In chlorine, the orbital accepting an electron to form chloride ion is 3p-orbital, which is away form the nucleus.

(v) Therefore, the electron-electron repulsion is less and more energy is liberated, when an electron is added to a chlorine atom forming a chloride anion. Thus, fluorine has less electron affinity than chlorine.

Example – 28

The first ($\Delta_i H_1$) and the second ($\Delta_i H_2$) ionization enthalpies (in kJ mol^{-1}) and the ($\Delta_{eg} H$) electron gain enthalpy (in kJ mol^{-1}) of a few elements are given below

Elements	**ΔH_1**	**ΔH_2**	**$\Delta_{eg} H$**
I	**520**	**7300**	**– 60**
II	**419**	**3051**	**– 48**
III	**1681**	**3374**	**– 328**
IV	**1008**	**1846**	**– 295**
V	**2372**	**5251**	**+ 98**
VI	**738**	**1451**	**– 40**

Which of the above elements is likely to be :

(a) The least reactive element

(b) The most reactive metal

(c) The most reactive non-metal

(d) The least reactive non-metal

(e) The metal which can form a stable binary halide of the formula MX_2 (X = halogen).

Sol. (a) The element V has highest first ionization enthalpy and positive electron gain enthalpy and hence it is likely to be the least reactive element.

(b) The element II which has the least first ionization enthalpy and a low negative electron gain enthalpy is the most reactive metal.

(c) The element III which has high first ionization enthalpy and a very high negative electron gain enthalpy is likely to be the most reactive non-metal.

(d) The element IV has a high negative electron gain enthalpy but not so high first ionization enthalpy. Therefore, it is the least reactive non-metal.

(e) The element VI has low values for first and second ionization enthalpies. Therefore, it appears that the element is an alkaline earth metal and hence will form binary halide of the formula MX_2.

Example – 29

Why are electron gain enthalpies of Be and Mg positive ?

Sol. They have fully filled s-orbitals and hence have no tendency to accept an additional electron. Consequently, energy has to be supplied if an extra electron has to be added to the much higher energy p-orbitals of the valence shell. That is why electron gain enthalpies of Be and Mg are positive.

Example – 30

Use the periodic table to answer the following questions :

(a) Identify an element with five electrons in the outer subshell.

(b) Identify the element that would tend to lose two electrons.

(c) Identify the element that would tend to gain two electons.

(d) Identify the group having metal, non-metal, liquid as well as gas at room temperature.

Sol. (a) The genral electronic configuration of the elements having five electrons in the outer subshell is $ns^2\ np^5$. This electronic configuration is characteristic of elements of group 17, i.e., halogens and their examples are F, Cl, Br, I, At, etc.

(b) The elements which have a tendency to lose two electrons must have two electrons in the valence shell. Therefore, their general electronic configuration should be $n\,s^2$. This electronic configuration is characteristic of group 2 elements, i.e., alkaline earth metals and their examples are Mg, Ca, Sr, Ba, etc.

(c) The elements which have a tendency to accept two electrons must have six electrons in the valence shell. Therefore, their general electronic configuration is $ns^2\ np^4$. This electronic configuration is characteristic of group 16 elements and their examples are O and S.

(d) A metal which is liquid at room temperature is mercury. It is a transition metal and belongs to group 12. A non-metal which is a gas at room temperature is hydrogen (group 1), nitrogen (group 15), oxygen (group 16), fluorine, chlorine (group 17) and inert gases (group 18).

A non-metal which is a liquid at room temperature is bromine (group 17).

Example – 31

What are major differences between metals and non-metals ?

Sol. Elements which have a strong tendency to **lose electrons** to form cations are called **metals** while those which have a strong tendency to **accept electrons** to form anions are called **non-metals**. Thus, metals are strong reducing agents, they have low ionization enthalpies, have less negative electron gain enthalpies, low electronegativity, form basic oxides and ionic compounds.

Non-metals, on the other hand, are strong oxidising agents, they have high ionization enthalpies, have high negative electron gain enthalpies, high electronegatively, form acidic oxides and covalent compounds.

Example – 32

Distinguish between electronegativity and electron affinity.

Sol.

	Electron affinity		Electronegativity
1.	Electron gain enthalpy or electron affinity is defined as the amount of energy released when neutral gaseous atom, accepts an electron to form an anion.	1.	Electronegativity of an atom in a molecule is defined as the tendency of an atom to attract towards itself the shared pair of electrons.
2.	It is expressed in eV/atom (electron volt per atom) or in kil number) joules per mol (kJ mol^{-1}).	2.	It does not have any unit (it is a number)
3.	It is a kind of absolute property of the elements.	3.	It is a relative term (atoms are compared with fluorine, whose assigned value of electronegativity is 4.0
4.	Electron affinity value is measured when the atoms are in their gaseous state.	4.	It is measured when the atoms are in their combined state (in state molecules).

Example – 33

Flourine is most electronegative element.

Sol. Electronegativity of an atom in a molecule is defined as the tendency of an atom to attract the shared pair of electrons towards itself.

In the periodic table, electronegativity increases with the increase in the atomic number across a period and decreases down the group.

Due to small atomic size of Fluorine, attraction between the nucleus of Fluorine and the shared pair of electron in a molecule is maximum.

Moreover, noble gas (Group - 18) have stable configuration and halides (Group - 17) are the most electronegative in a given period.

Therefore Fluorine is the most electronegative element.

Example – 34

Halogens except Flourine shows positive oxidation state of +1, +2, +3, +5, and +7.

Sol. (i) The outer Electronic configuration of halogens are ns^2, np^5 they can gain one electron and show a common oxidation state of –1.

(ii) The other halogen exhibit higher oxidation state as, +1, +2, +3, +5, and +7 due to vacant d-orbitals in their shell.

(iii) Since Flourine does not have d-orbital, it only exhibits only '–1' oxidation state.

Therefore Halogens except fluorine shows positive oxidation state +1, +2, +3, +5 and +7.

Example – 35

Considering the elements B, C, N, F and Si, the correct order of their non-metallic character is

(a) B > C > Si > N > F **(b) Si > C > B > N > F**

(c) F > N > C > B > Si **(d) F > N > C > Si > B**

Sol. In a period, the non-metallic character increases from left to right. Thus, among B, C, N and F, non-metallic character decreases in the order : F > N > C > B. However, within a group, non-metallic character decreases from top to bottom. Thus, C is more non-metallic than Si. Therefore, the correct sequence of decreasing non-metallic character is :

F > N > C > B > Si, i.e., option (c) is correct.

Example – 36

The increasing order of reactivity among group 1 elements is Li < Na < K < Rb < Cs whereas that of group 17 is F > Cl > Br > I. Explain.

Sol. The elements of group 1 have only one electron in their respective valence shells and thus have a strong tendency to lose this electron. The tendency to lose electrons, in turn, depends upon the ionization enthalpy. Since the ionization enthalpy decreases down the group, therefore, the reactivity of group 1 elements increases in the same order : Li < Na < K < Rb < Cs. In contrast, the elements of group 17, have seven electrons in their respective valence shells and thus have a strong tendency to accept one more electron. The tendency to accept electrons, in turn, depends upon their electrode potentials. Since the electrode potentials of group 17 elements decrease in the order : F (+ 2.87 V) > Cl (+ 1.36 V), Br (1.08 V) and I (+ 0.53 V), therefore, their **reactivities** also decrease in the same order : **F > Cl > Br > I**.

Alternatively, tendency to accept electrons can be linked to electron gain enthalpy. Since electron gain enthalpy becomes less and less negative as we move from Cl to I, therefore, reactivity increases from Cl to I. F is the **most reactive** due to its low bond dissociation energy.

Example – 37

Among alkali metals which element do you expect to be least electronegative and why ?

Sol. Electronegativity decreases as the size of the atom increases. Since Fr has the largest size, therefore, it has the least electronegativity.

Example – 38

How does the metallic and non metallic character vary on moving from left to right in a period ?

Sol. On moving from left to right in a period, the number of valence electrons increases by one at each succeeding element but the number of shells remains the same. As a result, the nuclear charge increases and the tendency of the element to lose electron decreases and hence the metallic character decreases as we move from left to right in a period. Conversely, as the nuclear charge increases, the tendency of the element to gain electrons increases and hence the non-metallic increases from left to right in a period.

Alternatively, metallic character decreases and non-metallic character increases as we move from left to right in a period. It is due to increase in ionization and electron gain enthalpy.

Example – 39

Explain 'Electron affinity'. Explain various factors affecting it ?

Sol. **(i) Def :**

Electron gain enthalpy or electron affinity is defined as the amount of energy released, when neutral gaseous atom, accepts an electron to form an anion.

Eg :-

$$Cl_{(g)} + e^{-} \longrightarrow Cl^{-1}_{(g)} + \underset{\left(\substack{\text{Corresponds to}\\ \text{Electron affinity}}\right)}{\text{Energy}}$$

(ii) **Unit :** ev/atom (electron volt per atom)

or

kJ/mol (kilo joules per mole of atom)

(iii) Greater the electron affinity, greater is the non-metallic character.

Factors affecting electron affinity :-

(i) **Atomic size :** Electron affinity increases with the decrease in atomic size

$$\boxed{\text{i.e. Electron affinity } \alpha \frac{1}{\text{Atomic size}}}$$

(ii) **Nuclear charge :** Electron affinity increases with the increase in the nuclear charge

$$\boxed{\text{i.e. Electron affinity } \alpha \text{ Nuclear charge}}$$

(iii) **Screening effect :** Electron affinity decreases with the increasing screeing affect (shielding effect) of electrons.

$$\boxed{\text{Electron affinity } \alpha \frac{1}{\text{Shielding effect}}}$$

(iv) **Electronic configuration :** Elements having stable electronic configuration shows poor tendency to accept an electron and hence electron affinity is less.

Example – 40

Give the formula of a species that will be isoelectronic with K^+ ion.

Sol. Isoelectronic species are those which have same number of electrons.

K^+ has 18 electrons.

$\therefore$ The species P^{3-}, S^{2-}, Cl^-, Ar, Ca^{2+} etc. are isoelectronic to K^+.

Example – 41

To which block (s, p, d or f) does the element with atomic number 50 belong ?

Sol. The electronic configuration of element with atomic number 50 is :

$1s^2$, $2s^22p^6$, $3s^23p^63d^{10}$, $4s^24p^64d^{10}$, $5s^25p^2$

The last electron enters into 5p-orbital. Hence, it is a p-block element.

Example – 42

What is the group number, period and block of the element with atomic number 43 ?

Sol. The electronic configuration of the element with atomic number 43 is

$1s^2$, $2s^22p^6$, $3s^23p^63d^{10}$, $4s^24p^64d^5$, $5s^2$

Since, the last electron is accommodated in d-subshell, the element belongs to d-block. The principal quantum number of outermost shell is 5, the element belongs to 5th period.

Group number of the element = 5 + 2 = 7 i.e.,

The element belongs to group 7.

Example – 43

Give reasons :

The ionic size of Cl^- ion is greater than K^+ ion, though both are isoelectronic.

Sol. This is because K^+ ion has greater nuclear charge (19) than that of Cl^- (17) and, thus, force of attraction towards nucleus is more in K^+ ion which brings contraction in size.

Example – 44

Which of the following species will have the largest and smallest size ?

Mg, Mg^{2+}, A*l*, Al^{3+}

Sol. Mg and A*l*, both belong to same period.

	Mg	A*l*
Atomic number	12 ;	13

Atomic size decreases from left to right across the period. Thus, Mg atom is larger in size that A*l* atom.

Cation is smaller than its neutral atom. Mg^{2+} ion is smaller than Mg atom and Al^{3+} ion is smaller than A*l* atom. Thus, Al^{3+} ion size is smallest and Mg atom is largest in size among the given species.

Example – 45

The first ionisation energy of carbon atom is greater than that of boron atom, whereas reverse is true for the second ionisation energy. Explain.

Sol. The electronic configurations of carbon and boron are as follows :

C : $1s^2, 2s^2\,2p_x^1\,2P_y^1$

B : $1s^2, 2s^2\,2p_x^1$

Due to higher nuclear charge in carbon, the force of attraction towards valency electron is more in carbon atom and hence the first ionisation energy is greater than boron atom. After loss of one electron, the monovalent cations have the configurations as follows :

B^+ : $1s^2, 2s^2$

C^+ : $1s^2, 2s^2\,2p_x^1$

The B^+ configuration is stable one and hence the removal of electron is difficult in comparison to C^+. Hence, second ionisation potential of boron is higher than carbon.

Example – 46

Why N has higher 1st ionisation potential than O-atom ?

Sol. The electronic configurations of nitrogen and oxygen are as follows :

N : $1s^2, 2s^2\,2p_x^1\,2p_y^1\,2p_z^1$

O : $1s^2, 2s^2\,2p_x^2\,2p_y^1\,2p_z^1$

In N, p-orbitals are half filled and hence, its electronic configuration is stable. It requires more energy to remove an electron. Hence, the IP of nitrogen is higher than oxygen atom which has less stable electronic configuration.

Example – 47

Why Mg has higher 1st ionisation potential than A*l*-atom ?

Sol. The electronic configurations of Mg and A*l* are as follows

Mg : $1s^2, 2s^2\,2p^6, 3s^2$

A*l* : $1s^2, 2s^2\,2p^6, 3s^2\,3p^1$

It is difficult to remove an electron from 3s in comparison to 3p (3s paired and 3p singly occupied). Hence, IP of Mg is higher than Al.

Example – 48

Out of Na^+ & Ne which has higher ionization enthalpy. Explain why.

Sol. Na^+ has higher ionization enthalpy than Ne. Na^+ & Ne are isoelectronic species. However, the nuclear charge in Na^+ is more than in Ne. Hence, the electrons are more tightly held in Na^+ & it has higher ionization enthalpy.

Example – 49

Why halogens have highest negative electron gain enthalpies in their respective periods ?

Sol. Because of small size and high effective nuclear charge.

Example – 50

Why chlorine has higher electron affinity than F ?

Sol. The size of fluorine atom is small and thus electron density is high. This resists the addition of electron, hence electron affinity of fluorine is less.

Example – 51

Which has the higher electronegativity : of N and F ?

Sol. Fluorine has a higher electronegativity since its size is smaller.

EXERCISE - 1 : BASIC OBJECTIVE QUESTIONS

Dobereiner's Law of Triads

1. Which of the following is not a Dobereiner triad ?

(a) Cl, Br, I (b) Ca, Sr, Ba

(c) Li, Na, K (d) Fe, Co, Ni

Modern Periodic Table/General Classification of Elements

2. Elements whose outer electronic configuration vary from ns^2np^1 to ns^2np^6 constitute

(a) s-Block of elements (b) p-Block of elements

(c) d-Block of elements (d) f-Block of elements

3. The basis of modern periodic table is

(a) atomic volume (b) atomic number

(c) atomic weights (d) atomic size

4. In the fourth period of the periodic table, how many elements have one or more 4d electrons ?

(a) 2 (b) 18

(c) 0 (d) 6

5. If the aufbau principle had not been followed, Ca (Z = 20) would have been placed in the :

(a) s-block (b) p-block

(c) d-block (d) f-block

6. La (lanthanum) having atomic number 57 is a member of :

(a) s-block elements (b) p-block elements

(c) d-block elements (d) f-block elements

7. Which of the following pairs has both members from the same group of periodic table

(a) Mg, Ba (b) Mg, Na

(c) Mg, Cu (d) Mg, Cl

8. The transition elements have a characteristic electronic configuration which can be represented as:

(a) $(n-2)\ s^2p^6d^{1-10}\ (n-1)\ s^2p^6ns^2$

(b) $(n-2)\ s^2p^6d^{1-10}\ (n-1)\ s^2p^6d^{1\text{ or }2}\ ns^1$

(c) $(n-1)s^2p^6d^{10}\ ns^2np^6nd^{1-10}$

(d) $(n-1)\ s^2p^6d^{1-10}ns^{0-2}$

9. The element whose electronic configuration is $1s^2, 2s^2, 2p^6, 3s^2$ is

(a) metal (b) metalloid

(c) inert gas (d) non – metal

10. Which of the following pairs do not show diagonal relationship ?

(a) Li and Mg (b) Be and Al

(c) B and Si (d) C and S

11. In which of the following pair, both the species are isoelectronic but first one is large in size than the second ?

(a) S^{2-}, O^{2-} (b) Cl^-, S^{2-}

(c) F^-, Na^+ (d) N^{3-}, P^{3-}

12. The correct order of ionic size of N^{3-}, Na^+, F^-, Mg^{2+} and O^{2-} is :

(a) $Mg^{2+} > Na^+ > F^- > O^{2-} < N^{3-}$

(b) $N^{3-} < F^- > O^{2-} > Na^+ > Mg^{2+}$

(c) $Mg^{2+} < Na^+ < F^- < O^{2-} < N^{3-}$

(d) $N^{3-} > O^{2-} > F^- > Na^+ < Mg^{2+}$

Prediction of Period, Group and Block

13. Ce (58) is a member of :

(a) s – block (b) p – block

(c) d – block (d) f – block

14. The electronic configuration of an element is $1s^2, 2s^22p^6, 3s^23p^3$. What is the atomic number of the element which is just below the above element in the periodic table.

(a) 34 (b) 49

(c) 33 (d) 31

Atomic Radius

15. Which of the following atom has largest size

(a) Cs (b) K

(c) Kr (d) Xe

16. Calculate the bond length of C-C bond if covalent radius of carbon is .77 Å

(a) .77 Å (b) 1.54 Å

(c) 1.86 Å (d) 1.29 Å

17. In comparison to the parent atom, the size of the
(a) Cation is smaller but anion is larger
(b) Cation is larger but anion is smaller
(c) Cation and anion are equal in size
(d) All the three are correct depending upon the atom

18. Which one is the correct order of the size of the iodine species.
(a) $I > I^+ > I^-$ (b) $I > I^- > I^+$
(c) $I^+ > I^- > I$ (d) $I^- > I > I^+$

19. Arrange the following elements in the order of increasing atomic size Cl, S, P, Ar
(a) Ar, Cl, S, P (b) Cl, S, P, Ar
(c) S, Cl, P, Ar (d) Ar, P, S, Cl

20. Which of the following ions has the smallest radius ?
(a) Li^+ (b) Na^+
(c) Be^{2+} (d) K^+

21. In iso – electronic species of Mg^{2+}, N^{3-}, Al^{3+}, the order of decreasing ionic radii will be
(a) $N^{3-} > Mg^{2+} > Al^{3+}$ (b) $Mg^{2+} > Al^{3+} > N^{3-}$
(c) $Al^{3+} > N^{3-} > Mg^{2+}$ (d) $Al^{3+} = Mg^{2+} < N^{3-}$

22. When a chlorine atom becomes chloride ion, its size
(a) remains unaltered (b) increases
(c) decreases (d) none of these

Ionization Potential or Ionization Energy

23. Lowest ionisation potential in periods is shown by :
(a) inert gases (b) halogens
(c) alkali metals (d) alkaline earth metals

24. The correct arrangement of the elements in the order of decreasing ionization energies is
(a) Na > Mg > Al (b) Mg > Na > Al
(c) Al > Mg > Na (d) Mg > Al > Na

25. The maximum tendency to form unipositive ion is for the element which has the following electronic configuration :
(a) $1s^2, 2s^2, 2p^6, 3s^2$ (b) $1s^2, 2s^2, 2p^6, 3s^2, 3p^1$
(c) $1s^2, 2s^2, 2p^6$ (d) $1s^2, 2s^2, 2p^6, 3s^2, 3p^3$

26. Which element has the highest ionisation energy ?
(a) Hydrogen (b) Lithium
(c) Boron (d) Sodium

27. An element will have lowest ionisation potential when its electronic configuration is
(a) $1s^1$ (b) $1s^2, 2s^2, 2p^2$
(c) $1s^2, 2s^2, 2p^5$ (d) $1s^2, 2s^2, 2p^6, 3s^1$

28. Which of the following iso – electronic ions has the lowest ionisation energy ?
(a) K^+ (b) Ca^{2+}
(c) Cl^{-1} (d) S^{2-}

29. The correct order of increasing ionisation potentials of K^+, Ar, Cl^- is
(a) $K^+ < Ar < Cl^-$ (b) $Cl^- < K^+ < Ar$
(c) $Cl^- < Ar < K^+$ (d) $Ar < Cl^- < K^+$

30. The first, second and third ionsiation energies (E_1, E_2 & E_3) for an element are 7 eV, 12.5 eV and 42.5 eV respectively. The most stable oxidation state of the element will be :
(a) +1 (b) +4
(c) +3 (d) +2

31. The order of ionisation potential between He^+ ion and H-atom (both species are in gaseous state) is :
(a) I.P. (He^+) = I.P. (H) (b) I.P. (He^+) < I.P. (H)
(c) I.P. (He^+) > I.P. (H) (d) cannot be compared

32. The first four I.E. values of an element are 284, 412, 656 and 3210 kJ mol^{-1}. The number of valence electrons in the element are :
(a) one (b) two
(c) three (d) four

33. The correct order of second I.E. of C, N, O and F are in the order :
(a) F > O > N > C (b) C > N > O > F
(c) O > N > F > C (d) O > F > N > C

34. The element which has highest first ionization energy in the periodic table is
(a) H (b) Rn
(c) F (d) He

35. Correct order of first ionization potential among the following elements Be, B, C, N, O is
(a) B < Be < C < O < N (b) B < Be < C < N < O
(c) Be < B < C < N < O (d) Be < B < C < O < N

36. The most electropositive element is :

(a) Cs (b) Ga

(c) Li (d) Pb

Electronegativity

37. Which of the following represent highly electropositive as well as highly electronegative element in its period

(a) Nitrogen (b) Fluorine

(c) Hydrogen (d) None

38. Outermost electronic configuration of least electronegative element in the periodic table is

(a) $2s^2\ 2p^5$ (b) $3s^2\ 3p^5$

(c) $2s^2\ 2p^4$ (d) $6s^2\ 6p^6\ 7s^1$

39. The element with highest electronegativity value is

(a) F (b) Cl

(c) P (d) N

40. With respect to chlorine, hydrogen will be

(a) Electropositive (b) Electronegative

(c) Neutral (d) None of these

41. Aqueous solutions of two compounds $M_1 - O - H$ and $M_2 - O - H$ are prepared in two different beakers. If, the electronegativity of $M_1 = 3.4$, $M_2 = 1.2$, $O = 3.5$ and $H = 2.1$, then the nature of two solutions will be respectively :

(a) acidic, basic (b) acidic, acidic

(c) basic, acidic (d) basic, basic

42. Which one of these is basic.

(a) CO_2 (b) SnO_2

(c) NO_2 (d) SO_2

43. The electronegativity of Cl, F, O, S increases in the order of

(a) S, O, Cl, F (b) S, Cl, O, F

(c) Cl, S, O, F (d) S, O, F, Cl

Electron Affinity

44. The correct order for electron affinities is a

(a) $F > Br > I$ (b) $F < Br < I$

(c) $F < I > Br$ (d) $Br < I < F$

45. Which of the following element is expected to have highest electron affinity

(a) $1s^2 2s^2 2p^6 3s^2 3p^5$ (b) $1s^2\ 2s^2\ 2p^3$

(c) $1s^2\ 2s^2\ 2p^4$ (d) $1s^2\ 2s^2\ 2p^5$

46. Which one of the following statements is incorrect ?

(a) Greater is the nuclear charge, greater is the electron gain enthalpy

(b) Nitrogen has almost zero electron gain enthalpy

(c) Electron gain enthalpy decreases from fluorine to iodine in the group

(d) Chlorine has highest electron gain enthalpy

47. The value of electron affinity for noble gases is likely to be

(a) high (b) low

(c) zero (d) positive

48. In which of the following processes, energy is liberated

(a) $Cl \rightarrow Cl^+ + e^-$ (b) $HCl \rightarrow H^+ + Cl^-$

(c) $O^- + e^- \rightarrow O^{2-}$ (d) $F + e^- \rightarrow F^-$

49. Which of the following process involves the gain of energy ?

(a) $O\,(g) + e^- \rightarrow O^-(g)$

(b) $Na^+ + e^- \rightarrow Na$

(c) $O^-\,(g) + e^- \rightarrow O^{2-}\,(g)$

(d) $O^{2-}\,(g) \rightarrow O^-(g) + e^-$

50. Second electron gain enthalpy :

(a) is always negative

(b) is always positive

(c) can be positive or negative

(d) is always zero

51. The correct order of increasing electron affinity of the following elements is :

(a) $O < S < F < Cl$ (b) $O < S < Cl < F$

(c) $S < O < F < Cl$ (d) $S < O < Cl < F$

Mixed

52. Which one of the following is incorrect ?

(a) An element which has high electronegativity always has high electron affinity

(b) Electron affinity is the property of an isolated atom

(c) Electronegativity is the property of a bonded atom

(d) Both electronegativity and electron affinity are usually directly related to nuclear charge and inversely related to atomic size.

53. Which of the following order is wrong.

(a) $NH_3 < PH_3 < AsH_3$ – Acidic

(b) $Li < Be < B < C$ – IE

(c) $Al_2O_3 < MgO < Na_2O < K_2O$ – Basic

(d) $Li^+ < Na^+ < K^+ < Cs^+$ – Ionic radius

EXERCISE - 2 : PREVIOUS YEAR JEE MAINS QUESTION

1. Ce^{3+}, La^{3+}, Pm^{3+} and Yb^{3+} have ionic radii in the increasing order as **(2002)**

(a) $La^{3+} < Ce^{3+} < Pm^{3+} < Yb^{3+}$

(b) $Yb^{3+} < Pm^{3+} < Ce^{3+} < La^{3+}$

(c) $La^{3+} = Ce^{3+} < Pm^{3+} < Yb^{3+}$

(d) $Yb^{3+} < Pm^{3+} < La^{3+} < Ce^{3+}$

2. According to the periodic law of elements, the variation in properties of elements is related to their **(2003)**

(a) atomic masses (b) nuclear masses

(c) atomic numbers

(d) nuclear neutron-proton number ratios

3. The radius of La^{3+} (atomic number : La = 57) is 1.06 Å. Which one of the following given values will be closest to the radius of Lu^{3+} (atomic number : Lu = 71) ? **(2003)**

(a) 1.60Å (b) 1.40 Å

(c) 1.06 Å (d) 0.85 Å

4. Which of the following groupings represents a collection of isoelectronic species ? **(2003)**

(At. Nos. Cs = 55, Br = 35)

(a) Ca^{2+}, Cs^{+}, Br (b) Na^{+}, Ca^{2+}, Mg^{2+}

(c) N^{3-}, F^{-}, Na^{+} (d) Be, Al^{3+}, Cl^{-}

5. The atomic numbers of vanadium (V), chromium (Cr), manganese (Mn) and iron (Fe) are respectively. 23, 24, 25 and 26. Which one of these may be expected to have the highest second ionization enthalpy ? **(2003)**

(a) Fe (b) V

(c) Cr (d) Mn

6. Which of the following ions has the highest value of ionic radius ? **(2004)**

(a) Li^{+} (b) B^{3+}

(c) O^{2-} (d) F^{-}

7. The formation of oxide ion, O^{2-} (g) requires first an exothermic and then an endothermic step as shown below

$O(g) + e^{-} \rightarrow O^{-}(g)$; $\Delta H^{o} = -142\ kJ\ mol^{-1}$

$O^{-}(g) + e^{-} \rightarrow O^{2-}(g)$; $\Delta H^{o} = 844\ kJ\ mol^{-1}$

This is because **(2004)**

(a) Oxygen is more electronegative

(b) Oxygen has high electron affinity

(c) O^{-} ion will tend to resist the addition of another electron

(d) O^{-} has comparatively larger size than oxygen atom

8. Which one of the following sets of ions represents the collection of isoelectronic species ? **(2004)**

(a) K^{+}, Ca^{2+}, Sc^{3+}, Cl^{-} (b) Na^{+}, Ca^{2+}, Sc^{3+}, F^{-}

(c) K^{+}, Cl^{-}, Mg^{2+}, Sc^{3+} (d) Na^{+}, Mg^{2+}, Al^{3+}, Cl^{-}

9. Among Al_2O_3, SiO_2, P_2O_3 and SO_2 the correct order of acid strength is **(2004)**

(a) $SO_2 < P_2O_3 < SiO_2 < Al_2O_3$

(b) $SiO_2 < SO_2 < Al_2O_3 < P_2O_3$

(c) $Al_2O_3 < SiO_2 < SO_2 < P_2O_3$

(d) $Al_2O_3 < SiO_2 < P_2O_3 < SO_2$

10. In which of the following arrangements the order is not according to the property indicated against it ? **(2005)**

(a) Li < Na < K < Rb : Increasing metallic radius

(b) I < Br < F < Cl : Increasing electron gain enthalpy (with negative sign)

(c) B < C < N < O : Increasing first ionization enthalpy

(d) $Al^{3+} < Mg^{2+} < Na^{+} < F^{-}$: Increasing ionic size.

11. Which of the following sets of ions represents a collection of isoelectronic species ? **(2006)**

(a) N^{3-}, O^{2-}, F^{-}, S^{2-} (b) Li^{+}, Na^{+}, Mg^{2+}, Ca^{2+}

(c) K^{+}, Cl^{-}, Ca^{2+}, Sc^{3+} (d) Ba^{2+}, Sr^{2+}, K^{+}, Ca^{2+}

12. Following statements regarding the periodic trends of chemical reactivity of the alkali metals and the halogens are given. Which of these statements give the correct picture ? **(2006)**

(a) The reactivity decreases in the alkali metals but increases in the halogens with increase in atomic number down the group

(b) In both the alkali metals and the halogens the chemical reactivity decreases with increase in atomic number down the group

(c) Chemical reactivity increase with increase in atomic number down the group in both the alkali metals and halogens.

(d) In alkali metals the reactivity increases but in the halogens it decreases with increase in atomic number down the group

13. The increasing order of the first ionization enthalpies of the elements B, P, S and F (lowest first) is **(2007)**

(a) $F < S < P < B$ (b) $P < S < B < F$

(c) $B < P < S < F$ (d) $B < S < P < F$

14. The set representing the correct order of ionic radius is **(2009)**

(a) $Li^+ > Be^{2+} > Na^+ > Mg^{2+}$

(b) $Na^+ > Li^+ > Mg^{2+} > Be^{2+}$

(c) $Li^{2+} > Na^+ > Mg^{2+} > Be^{2+}$

(d) $Mg^{2+} > Be^{2+} > Li^+ > Na^+$

15. The correct sequence which shows decreasing order of the ionic radii of the elements is **(2010)**

(a) $Al^{3+} > Mg^{2+} > Na^+ > F^- > O^{2-}$

(b) $Na^+ > Mg^{2+} > Al^{3+} > O^{2-} > F^-$

(c) $Na^+ > F^- > Mg^{2+} > O^{2-} > Al^{3+}$

(d) $O^{2-} > F^- > Na^+ > Mg^{2+} > Al^{3+}$

16. Which one of the following orders presents the correct sequence of the increasing basic nature of the given oxides ? **(2011)**

(a) $Al_2O_3 < MgO < Na_2O < K_2O$

(b) $MgO < K_2O < Al_2O_3 < Na_2O$

(c) $Na_2O < K_2O < MgO < Al_2O_3$

(d) $K_2O < Na_2O < Al_2O_3 < MgO$

17. The correct order of electron gain enthalpy with negative sign of F, Cl, Br and I, having atomic number 9, 17, 35 and 53 respectively, is **(2011)**

(a) $I > Br > Cl > F$ (b) $F > Cl > Br > I$

(c) $Cl > F > Br > I$ (d) $Br > Cl > I > F$

18. The increasing order of the ionic radii of the given isoelectronic species is **(2012)**

(a) $Cl^-, Ca^{2+}, K^+, S^{2-}$ (b) $S^{2-}, Cl^-, Ca^{2+}, K^+$

(c) $Ca^{2+}, K^+, Cl^-, S^{2-}$ (d) $K^+, S^{2-}, Ca^{2+}, Cl^-$

19. Which of the following represents the correct order of increasing first ionization enthalpy for Ca, Ba, S, Se and Ar ? **(2013)**

(a) $Ca < S < Ba < Se < Ar$ (b) $S < Se < Ca < Ba < Ar$

(c) $Ba < Ca < Se < S < Ar$ (d) $Ca < Ba < S < Se < Ar$

20. The first ionisation potential of Na is 5.1 eV. The value of electron gain enthalpy of Na^+ will be **(2013)**

(a) – 2.55 eV (b) – 5.1 eV

(c) – 10.2 eV (d) + 2.55 eV

21. The ionic radii (in Å) of N^{3-}, O^{2-} and F^- are respectively : **(2015)**

(a) 1.71, 1.40 and 1.36 (b) 1.71, 1.36 and 1.40

(c) 1.36, 1.40 and 1.71 (d) 1.36, 1.71 and 1.40

22. Which of the following atoms has the highest first ionization energy **(2016)**

(a) Na (b) K

(c) Sc (d) Rb

JEE MAINS ONLING QUESTION

1. Which of the following series correctly represents relations between the elements from X to Y? **(Online 2014 SET-2)**

$X \rightarrow Y$

(a) ${}_6C \rightarrow {}_{32}Ge$ Atomic radii increases

(b) ${}_9F \rightarrow {}_{35}Br$ Electron gain enthalpy with negative sign increases

(c) ${}_3Li \rightarrow {}_{19}K$ Ionization enthalpy increases

(d) ${}_{18}Ar \rightarrow {}_{54}Xe$ Noble character increases

2. Similarity in chemical properties of the atoms of elements in a group of the Periodic table is most closely related to: **(Online 2014 SET-3)**

(a) number of principal energy levels

(b) number of valence electrons

(c) atomic numbers

(d) atomic masses

3. Which of the following arrangements represents the increasing order (smallest to largest) of ionic radii of the given species O^{2-}, S^{2-}, N^{3-}, P^{3-} **(Online 2014 SET-3)**

(a) $O^{2-} < N^{3-} < S^{2-} < P^{3-}$ (b) $N^{3-} < O^{2-} < P^{3-} < S^{2-}$

(c) $N^{3-} < S^{2-} < O^{2-} < P^{3-}$ (d) $O^{2-} < P^{3-} < N^{3-} < S^{2-}$

4. Which one of the following has largest ionic radius ?

(Online 2014 SET-4)

(a) O_2^{2-} (b) Li^+

(c) F^- (d) B^{3+}

5. Which of the following represents the correct order of increasing first ionization enthalpy for Ca, Ba, S, Se and Ar? **(Online 2015 SET-1)**

(a) $S < Se < Ca < Ba < Ar$

(b) $Ba < Ca < Se < S < Ar$

(c) $Ca < Ba < S < Se < Ar$

(d) $Ca < S < Ba < Se < Ar$

6. Identify the correct order of the size of the following :

(Online 2016 SET-1)

(a) $Ca^{2+} < K^+ < Ar < S^{2-} < Cl^-$

(b) $Ca^{2+} < K^+ < Ar < Cl^- < S^{2-}$

(c) $Ar < Ca^{2+} < K^+ < Cl^- < S^{2-}$

(d) $Ca^{2+} < Ar < K^+ < Cl^- < S^{2-}$

7. The following statements concern elements in the periodic table. Which of the following is true?

(Online 2016 SET-2)

(a) All the elements in Group 17 are gases.

(b) The Group 13 elements are all metals.

(c) Elements of Group 16 have lower ionization enthalpy values compared to those of Group 15 in the corresponding periods.

(d) For Group 15 elements, the stability of +5 oxidation state increases down the group.

8. Consider the following ionization enthalpies of two elements 'A' and 'B'

(Online 2017 SET-1)

Element	Ionization enthalpy (kJ/mol)		
	1^{st}	2^{nd}	3^{rd}
A	899	1757	14847
B	737	1450	7731

Which of the following statements is **correct** ?

(a) Both 'A' and 'B' belong to group-1 where 'B' comes below 'A'.

(b) Both 'A' and 'B' belong to group-1 where 'A' comes below 'B'.

(c) Both 'A' and 'B' belong to group-2 where 'B' comes below 'A'.

(d) Both 'A' and 'B' belong to group-2 where 'A' comes below 'B'.

9. The electronic configuration with the highest ionization enthalpy is : **(Online 2017 SET-2)**

(a) $[Ne]\ 3s^2\ 3p^1$ (b) $[Ne]\ 3s^2\ 3p^2$

(c) $[Ne]\ 3s^2\ 3p^3$ (d) $[Ar]\ 3d^{10}\ 4s^2\ 4p^3$

10. Which one of the following is an acidic oxide ?

(Online 2017 SET-2)

(a) KO_2 (b) BaO_2

(c) SiO_2 (d) CsO_2

11. For Na^+, Mg^{2+}, F^- and O^{2-}; the correct order of increasing ionic radii is : **(Online 2018 SET-1)**

(a) $O^{2-} < F^- < Na^+ < Mg^{2+}$

(b) $Na^+ < Mg^{2+} < F^- < O^{2-}$

(c) $Mg^{2+} < Na^+ < F^- < O^{2-}$

(d) $Mg^{2+} < O^{2-} < Na^+ < F^-$

12. The correct order of electron affinity is :

(Online 2018 SET-2)

(a) F>Cl>O (b) F>O>Cl

(c) Cl>F>O (d) O>F>Cl

EXERCISE - 3 : ADVANCED OBJECTIVE QUESTIONS

1. All questions marked **"S"** are single choice questions
2. All questions marked **"M"** are multiple choice questions
3. All questions marked **"C"** are comprehension based questions
4. All questions marked **"A"** are assertion–reason type questions

   (A) If both assertion and reason are correct and reason is the correct explanation of assertion.

   (B) If both assertion and reason are true but reason is not the correct explanation of assertion.

   (C) If assertion is true but reason is false.

   (D) If reason is true but assertion is false.
5. All questions marked **"X"** are matrix–match type questions

Screening Effect

1. (S) The order of screening effect of electrons of s, p, d and f orbitals of a given shell of an atom on its outer shell electrons is:

(a) $s > p > d > f$ (b) $f > d > p > s$

(c) $p < d < s < f$ (d) $f > p > s > d$

Comprehension

Nuclear charge actually experienced by an electron is termed as effective nuclear charge. The effective nuclear charge Z* actually depends on type of shell and orbital in which electron is actually present. The relative extent to which the various orbitals penetrate the electron clouds of other orbitals is.

$s > p > d > f$ (for the same value of n)

The phenomenon in which penultimate shell electrons act as screen or shield in between nucleus and valence shell electrons and thereby reducing nuclear charge is known as shielding effect. The penultimate shell electrons repel the valence shell electron to keep them loosely held with nucleus. It is thus evident that more is the shielding effect, lesser is the effective nuclear charge and lesser is the ionization energy.

2. (C) Which of the following valence electron experience maximum effective nuclear charge ?

(a) $4s^1$ (b) $4p^1$

(c) $3d^1$ (d) $2p^3$

3. (C) Which of the following is not concerned to effective nuclear charge ?

(a) Higher ionization potential of carbon than boron

(b) Higher ionization potential of magnesium than aluminium

(c) Higher values of successive ionization energy

(d) Higher electronegativity of higher oxidation state

4. (C) Ionization energy is not influenced by :

(a) Size of atom

(b) Effective nuclear charge

(c) Electrons present in inner shell

(d) Change in entropy

Modern Periodic Table

5. (S) The electronic configuration of gadolinium (Atomic number 64) is

(a) $[Xe]\,4f^3\,5d^5\,6s^2$ (b) $[Xe]\,4f^7\,5d^2\,6s^1$

(c) $[Xe]\,4f^7\,5d^1\,6s^2$ (d) $[Xe]\,4f^8\,5d^6\,6s^2$

6. (S) The statement that is not correct for periodic classification of elements is :

(a) The properties of elements are periodic function of their atomic numbers.

(b) Non metallic elements are less in number then metallic elements.

(c) For transition elements, the 3d-orbitals are filled with electrons after 3p-orbitals and before 4s-orbitals.

(d) The first ionisation enthalpies of elements generally increase with increase in atomic number as we go along a period.

7. (S) The period number in the long form of the periodic table is equal to

(a) Magnetic quantum number of any element of the period.

(b) Atomic number of any element of the period.

(c) Maximum Principal quantum number of any element of the period.

(d) Maximum Azimuthal quantum number of any element of the period.

8. (A) **Assertion :** The 5th period of periodic table contains 18 elements not 32.

Reason : $n = 5, l = 0, 1, 2, 3, 4$. The order in which the energy of available orbitals 4d, 5s and 5p increases is 5s < 4d < 5p and the total number of orbitals available are 9 and thus 18 electrons can be accomodated.

(a) A (b) B

(c) C (d) D

General Classification of Elements

9. (S) The elements in which electrons are progressively filled in 4*f*-orbital are called

(a) Actinoids (b) Transition elements

(c) Lanthanoids (d) Halogens

10. (S) La (lanthanum) having atomic number 57 is a member of :

(a) s-block elements (b) p-block elements

(c) d-block elements (d) f-block elements

11. (S) If the aufbau principle had not been followed, Ca (Z = 20) would have been placed in the :

(a) s-block (b) p-block

(c) d-block (d) f-block

12. (A) **Assertion :** The 4f-and 5f-inner transition series of elements are placed separately at the bottom of the periodic table.

Reason : (i) This prevents the undue expansion of the periodic table i.e., maintains its structure.
(ii) This preserve the principles of classification by keeping elements with similar properties in a single column.

(a) A (b) B

(c) C (d) D

13. (M) Which of the following sets contain only isoelectronic ions?

(a) $Zn^{2+}, Ca^{2+}, Ga^{3+}, Al^{3+}$

(b) $K^{+}, Ca^{2+}, Sc^{3+}, Cl^{-}$

(c) $P^{3-}, S^{2-}, Cl^{-}, K^{+}$

(d) $Ti^{4+}, Ar, Cr^{3+}, V^{5+}$

Atomic Radius

14. (S) In which of the following sets of elements, they have nearly the same atomic size

(a) Li, Be, B, C (b) Mg, Ca, Sr, Ba

(c) O, S, Se, Te (d) Fe, Co, Ni, Cu

15. (S) The correct order of the sizes of C, N, P, S, is

(a) N < C < P < S (b) C < N < S < P

(c) C < N < P < S (d) N < C < S < P

16. (S) Consider the isoelectronic species, Na^{+}, Mg^{2+}, F^{-} and O^{2-}. The correct order of increasing length of their radii is ....

(a) $F^{-} < O^{2-} < Mg^{2+} < Na^{+}$

(b) $Mg^{2+} < Na^{+} < F < O^{2-}$

(c) $O^{2-} < F < Na^{+} < Mg^{2+}$

(d) $O^{2-} < F^{-} < Mg^{2+} < Na^{+}$

17. (S) In which of the following pair, both the species are isoelectronic but the first one is large in size than the second ?

(a) S^{2-}, O^{2-} (b) Cl^{-}, S^{2-}

(c) F^{-}, Na^{+} (d) N^{3-}, P^{3-}

18. (S) The correct order of ionic size of N^{3-}, Na^{+}, F^{-}, Mg^{2+} and O^{2-} is :

(a) $Mg^{2+} > Na^{+} > F^{-} > O^{2-} < N^{3-}$

(b) $N^{3-} < F^{-} > O^{2-} > Na^{+} > Mg^{2+}$

(c) $Mg^{2+} < Na^{+} < F^{-} < O^{2-} < N^{3-}$

(d) $N^{3-} > O^{2-} > F^{-} > Na^{+} < Mg^{2+}$

Ionization Potential or Ionization Energy

19. (S) From the ground state electronic configuration of the elements given below, pick up the one with highest value of second ionization energy

(a) $1s^2 2s^2 2p^6 3s^2$ (b) $1s^2 2s^2 2p^6 3s^1$

(c) $1s^2 2s^2 2p^6$ (d) $1s^2 2s^2 2p^5$

20. (S) In which of the following pairs, the ionization energy of the first species is less than that of the second

(a) N, P (b) Be^+, Be

(c) N, N^- (d) S, P

21. (S) IE_1 and IE_2 of Mg are 178 and 348 kcal mol^{-1} respectively. The energy required for the reaction Mg (g) $\rightarrow$ Mg^{2+} (g) + $2e^-$ is

(a) + 170 kcal (b) + 526 kcal

(c) – 170 kcal (d) – 525 kcal

22. (S) For a given value of n (principal quantum number), ionization energy is highest for

(a) d-Electrons (b) f-Electrons

(c) p-Electrons (d) s-Electrons

23. (S) Which one of the following is true about metallic character when the move from left to right in a period and top to bottom in a group

(a) Increases both in the period and group

(b) Decreases both in the period and group

(c) Decreases in the period and increases in a group

(d) Increases in the period and decreases in a group

24. (S) The electronic configurations of the elements X, Y, Z and J are given below. Which element has the highest metallic character ?

(a) X = 2, 8, 4 (b) Y = 2, 8, 8

(c) Z = 2, 8, 8, 1 (d) J = 2, 8, 8, 7

25. (S) The first ionisation enthalpies of Na, Mg, A*l* and Si are in the order:

(a) Na < Mg > A*l* < Si (b) Na > Mg > A*l* > Si

(c) Na < Mg < A*l* < Si (d) Na > Mg > A*l* < Si

26. (S) The ionization energy of boron is less than that of beryllium because :

(a) beryllium has a higher nuclear charge tha boron

(b) beryllium has a lower nuclear charge than boron

(c) the outermost electron in boron occupies a 2p-orbital

(d) the 2s and 2p-orbitials of boron are degenerate

27. (S) Which of the following is correct order of Ist, IInd and IIIrd ionization energy of nitrogen ?

(a) I > II > III (b) III > II > I

(c) I > II < III (d) II > I > III

28. (S) The graph of ionization energy and atomic number does not increase smoothly because

(a) the values for Be and Mg are high and this is attributed to the stability of a filled s level

(b) the values for N and P are high due to half filled p-orbital

(c) both (a) and (b) are correct

(d) None of these

29. (M) The first ionisation energy of first atom is greater than that of second atom, whereas reverse order is true for their second ionisation energy. Which set of elements is in accordance to above statement ?

(a) C > B (b) P > S

(c) Be > B (d) Mg > Na

Comprehension

The $(IE)_1$ and the $(IE)_2$ in kJ mol^{-1} of a few elements designated by Roman numerals are shown below :

Element	$(IE)_1$	$(IE)_2$
A	2372	5251
B	520	7300
C	900	1760
D	1680	3380

30. (C) Which of the above elements is likely to be a reactive metal ?

(a) A (b) B

(c) C (d) D

31. (C) Which of the above elements is likely to be a reactive non-metal ?

(a) A (b) B

(c) C (d) D

32. (C) Which represents a noble gas ?

(a) A (b) B

(c) C (d) D

33. (A) **Assertion :** Third ionisation energy of phosphorous is larger than sulphur.

Reason : There is a larger amount of stability associated with filled s- and p- sub-shells (a noble gas electron configuration) which corresponds to having eight electrons in the valence shell of an atom or ion.

(a) A (b) B

(c) C (d) D

Electronegativity

34. (S) Electronegativity values for the elements help in predicting

(a) Polarity of bonds

(b) Dipole moments

(c) Valency of elements

(d) Position in the electrochemical series

35. (S) Which one of the following elements shows both positive and negative oxidation states ?

(a) Cesium (b) Fluorine

(c) Iodine (d) Xenon

36. (S) The electronegativity of the following elements increases in the order :

(a) $C < N < Si < P$ (b) $Si < P < C < N$

(c) $N < C < P < Si$ (d) $C < Si < N < P$

37. (M) Which of the following effects the electronegativity of an atom ?

(a) s-Character in hybridization

(b) Multiplicity of bond between atoms

(c) Oxidation number

(d) The number of neutrons in the nucleus

38. (S) A, B and C are hydroxy-compounds of the elements X, Y and Z respectively. X, Y and Z are in the same period of the periodic table. A gives an aqueous solution of pH less than seven. B reacts with both strong acids and strong alkalis. C gives an aqueous solution which is strongly alkaline.

Which of the following statements is/are true ?

I : The three elements are metals.

II : The electronegativities decrease from X to Y to Z.

III : The atomic radius decreases in the order X, Y and Z.

IV : X, Y and Z could be phosphorus, aluminium and sodium respectively.

(a) I, II, III only correct (b) I, III only correct

(c) II, IV only correct (d) II, III, IV only correct

Electron Affinity/Electron Gain Enthalpy

39. (S) The electron affinities of B, C, N and O are in the order of

(a) $B < C < N < O$ (b) $B < C < O > N$

(c) $B < C > O > N$ (d) $B > C < O < N$

40. (S) Which one of the following statements is incorrect ?

(a) Greater is the nuclear charge, greater is the electron affinity

(b) Nitrogen has zero electron affinity

(c) Electron affinity decreases from fluorine to iodine in the group

(d) Chlorine has highest electron affinity.

41. (S) The lower electron affinity of fluorine than that of chlorine is due to

(a) Smaller size

(b) Smaller nuclear charge

(c) Difference in their electronic configurations

(d) Its highest reactivity

42. (S) Among halogens, the correct order of amount of energy released in electron gain (electron gain enthalpy) is:

(a) $F > Cl > Br > I$ (b) $F < Cl < Br < I$

(c) $F < Cl > Br > I$ (d) $F < Cl < Br < I$

43. (S) The formation of the oxide ion, $O^{2-}(g)$, from oxygen atom requires first an exothermic and then an endothermic step as shown below:

$O(g) + e^- \rightarrow O^-(g)\,;\,\Delta H^- = -141\ kJ\ mol^{-1}$

$O^-(g) + e^- \rightarrow O^{2-}(g)\,;\,\Delta H^- = +780\ kJ\ mol^{-1}$

Thus process of formation of O^{2-} in gas phase is unfavourable even though O^{2-} is isoelectronic with neon. It is due to the fact that,

(a) Oxygen is more electronegative.

(b) Addition of electron in oxygen results in larger size of the ion.

(c) Electron repulsion outweighs the stability gained by achieving noble gas configuration.

(d) O^- ion has comparatively smaller size than oxygen atom.

44. (S) The correct order of increasing electron affinity of the following elements is :

(a) $O < S < F < Cl$ (b) $O < S < Cl < F$

(c) $S < O < F < Cl$ (d) $S < O < Cl < F$

45. (M) Which of the following elements will gain one electron more readily in comparison to other elements of their group?

(a) S (g) (b) Na (g)

(c) O (g) (d) *Cl* (g)

46. (X) Electronic configuration of some elements is given in Column I and their electron gain enthalpies are given in Column II. Match the electronic configuration with electron gain enthalpy.

Column (I) **Electronic configuration**	**Column (II)** **Electron Gain Enthalpy/kJ mol^{-1}**
(A) $1s^2\ 2s^2\ sp^6$	(P) - 53
(B) $1s^2\ 2s^1\ 2p^6\ 3s^1$	(Q) - 328
(C) $1s^2\ 2s^2\ 2p^5$	(R) - 141
(D) $1s^2\ 2s^2\ 2p^4$	(S) + 48

Mixed

47. (S) Which of the following statements is/are wrong ?

(a) Van der Waals' radius of iodine is more than its covalent radius

(b) All isoelectronic ions belong to same period of the periodic table

(c) I.E.$_1$ of N is higher than that of O while I.E.$_2$ of O is higher than that of N

(d) The electron gain enthalpy of N is almost zero while that of P is 74.3 kJ mol^{-1}

48. (X) Match the correct ionisation enthalpes and electron gain enthaopies of the following elements.

Elements		ΔH_1	ΔH_2	$\Delta_{eg}H$
(A) Most reactive non metal	(P)	419	3051	- 48
(B) Most reactive metal	(Q)	1681	3374	- 328
(C) Least reactive element	(R)	738	1451	- 40
(D) Metal forming binary halide	(S)	2372	5251	+ 48

49. (S) Consider the following statements :

(I) The radius of an anion is larger than that of the parent atom.

(II) The ionization energy generally increases with increasing atomic number in a period.

(III) The electronegativity of an element is the tendency of an isolated atom to attract an electron.

Which of the above statements is/are correct ?

(a) I alone (b) II alone

(c) I and II (d) II and III

50. (M) Which of the following statements are correct?

(a) Helium has the highest first ionisation enthalpy in the periodic table.

(b) Chlorine has less negative electron gain enthalpy than fluorine.

(c) Mercury and bromine are liquids at room temperature.

(d) In any period, atomic radius of alkali metal is the highest.

EXERCISE - 4 : PREVIOUS YEAR JEE ADVANCED QUESTIONS

1. The increasing order of atomic radii of the following Group 13 elements is **(2016)**

(a) $Al < Ga < In < Tl$ (b) $Ga < Al < In < Tl$

(c) $Al < In < Ga < Tl$ (d) $Al < Ga < Tl < In$

2. Identify the least stable ion amongst the following **(2002)**

(a) Li^+ (b) Be^-

(c) B^- (d) C^-

3. The set representing the correct order of first ionization potential is **(2001)**

(a) $K > Na > Li$ (b) $Be > Mg > Ca$

(c) $B > C > N$ (d) $Ge > Si > C$

4. The correct order of radii is **(2000)**

(a) $N < Be < B$ (b) $F^- < O^{2-} < N^{3-}$

(c) $Na < Li < K$ (d) $Fe^{3+} < Fe^{2+} < Fe^{4+}$

5. The incorrect statement among the following is **(1997)**

(a) the first ionisation potential of Al is less than the first ionisation potential of Mg

(b) the second ionisation potential of Mg is greater than the second ionisation potential of Na

(c) the first ionisation potential of Na is less than the first ionisation potential of Mg

(d) the third ionisation potential of Mg is greater than third ionisation potential of Al

6. Which of the following has the maximum number of unpaired electrons ? **(1996)**

(a) Mg^{2+} (b) Ti^{3+}

(c) V^{3+} (d) Fe^{2+}

7. Which has most stable +2 oxidation state ? **(1995)**

(a) Sn (b) Pb

(c) Fe (d) Ag

8. The statement that is not correct for the periodic classification of element, is **(1992)**

(a) The properties of elements are the periodic functions of their atomic numbers.

(b) Non-metallic elements are lesser in number than metallic elements.

(c) The first ionization energies of elements along a period do not vary in a regular manner with increase in atomic number.

(d) For transition elements the d-subshells are filled with electrons monotonically with increase in atomic number.

9. Amongst the following elements (whose electronic configurations are given below), the one having the highest ionisation energy is **(1990)**

(a) $[Ne]\ 3s^2 3p^1$ (b) $[Ne]\ 3s^2 3p^3$

(c) $[Ne]\ 3s^2 3p^2$ (d) $[Ne]\ 3d^{10} 4s^2 4p^3$

10. Which one of the following is the smallest in size ? **(1989)**

(a) N^{3-} (b) O^{2-}

(c) F^- (d) Na^+

11. The electronegativity of the following elements increase in the order **(1987)**

(a) C, N, Si, P (b) N, Si, C, P

(c) Si, P, C, N (d) P, Si, N, C

12. The first ionisation potential electron volts of nitrogen and oxygen atoms are respectively given by **(1987)**

(a) 14.6, 13.6 (b) 13.6, 14.6

(c) 13.6, 13.6 (d) 14.6, 14.6

13. Atomic radii of fluorine and neon in angstrom units are respectively given by **(1987)**

(a) 0.72, 1.60 (b) 1.60, 1.60

(c) 0.72, 0.72 (d) None of these

14. The hydration energy of Mg^{2+} is larger than that of **(1984)**

(a) Al^{3+} (b) Na^+

(c) Be^{2+} (d) Mg^{3+}

15. The element with the highest first ionisation potential is **(1982)**

(a) boron (b) carbon

(c) nitrogen (d) oxygen

16. The correct order of second ionisation potential of carbon, nitrogen, oxygen and fluorine is **(1981)**

(a) $C > N > O > F$ (b) $O > N > F > C$

(c) $O > F > N > C$ (d) $F > O > N > C$

Objective Questions
(One or more than one correct option)

17. The option(s) with only amphoteric oxides is (are) **(2017)**

(a) NO, B_2O_3, PbO, SnO_2

(b) Cr_2O_3, CrO, SnO, PbO

(c) Cr_2O_3, BeO, SnO, SnO_2

(d) ZnO, Al_2O_3, PbO, PbO_2

18. Ionic radii of **(1999)**

(a) $Ti^{4+} < Mn^{7+}$ (b) $^{35}Cl^- < {}^{37}Cl^-$

(c) $K^+ > Cl^-$ (d) $P^{3+} > P^{5+}$

19. The first ionization potential of nitrogen and oxygen atoms are related as follows **(1989)**

(a) The ionization potential of oxygen is less than the ionization potential of nitrogen.

(b) The ionization potential of nitrogen is greater than the ionization potential of oxygen.

(c) The two ionization potential values are comparable.

(d) The difference between the two ionization potential is too large.

20. Sodium sulphate is soluble in water whereas barium sulphate is sparingly soluble because **(1989)**

(a) the hydration energy of sodium sulphate is more than its lattice energy

(b) the lattice energy of barium sulphate is more than its hydration energy

(c) the lattice energy has not role to play in solubility

(d) the hydration energy of sodium sulphate is less than its lattice energy

21. The statements that is/are true for the long form of the Periodic Table is/are **(1988)**

(a) If reflects the sequence of filling the electrons in the order of sub-energy level s, p, d and f

(b) It helps to predict the stable valency states of the elements

(c) If reflects tends in physical and chemical properties of the elements

(d) It helps to predict the relative ionicity of the bond between any two elements

Assertion and Reason

(A) If both assertion and reason are correct and reason is the correct explanation of assertion.

(B) If both assertion and reason are true but reason is not the correct explanation of assertion.

(C) If assertion is true but reason is false.

(D) If reason is true but assertion is false.

22. **Assertion :** The first ionization energy of Be is greater than that of B.

Reason : 2p orbital is lower in energy than 2s. **(2000)**

(a) A (b) B

(c) C (d) D

23. **Assertion :** F atom has a less negative electron affinity than Cl atom.

Reason : Additional electrons are repelled more effectively by 3p electrons in Cl atom than by 2p electrons in F atom. **(1998)**

(a) A (b) B

(c) C (d) D

Fill in the Blanks

24. Compounds that formally contain Pb^{4+} are easily reduced to Pb^{2+}. The stability of the lower oxidation state is due to .... **(1997)**

25. Ca^{2+} has a smaller ionic radius than K^+ because it has .... **(1993)**

26. On Mulliken scale, the average of ionization potential and electron affinity is known as .... **(1985)**

27. The energy released when an electron is added to a neutral gaseous atom is called .... **(1982)**

True/False

28. The decreasing order of electron affinity of F, Cl, Br is $F > Cl > Br$. **(1993)**

29. In group IA of alkali metals, the ionization potential decreases down the group. Therefore, lithium is a poor reducing agent. **(1987)**

Subjective Problems

30. The Periodic Table consists of 18 groups. An isotope of copper, on bombardment with protons, undergoes a nuclear reaction yielding element X as shown below. To which group, element X belongs in the Periodic Table ? **(2012)**

$$^{63}_{29}Cu + {}^{1}_{1}H \longrightarrow 6\,{}^{1}_{0}n + {}^{4}_{2}\alpha + 2\,{}^{1}_{1}H + X$$

31. Compare qualitatively the first and second ionization potentials of copper and zinc. Explain the observation. **(1996)**

32. Arrange the following in

(i) Decreasing ionic size : Mg^{2+}, O^{2-}, Na^{+}, F^{-} **(1985)**

(ii) Increasing acidic property : ZnO, Na_2O_2, P_2O_5, MgO **(1985)**

(iii) Increasing first ionisation potential : Mg, Al, Si, Na **(1985)**

(iv) Increasing size : Cl^{-}, S^{2-}, Ca^{2+}, Ar **(1986)**

(v) Increasing order of ionic size : N^{3-}, Na^{+}, F^{-}, O^{2-}, Mg^{2+} **(1991)**

(vi) Increasing order of basic character : MgO, SrO, K_2O, NiO, Cs_2O **(1991)**

(vii) Arrange the following ions in order of their increasing radii : Li^{+}, Mg^{2+}, K^{+}, Al^{3+}. **(1991)**

33. The first ionisation energy of carbon atom is greater than that of boron atom whereas, the reverse is true for the second ionisation energy. Explain. **(1989)**

ANSWER KEY

Exercise - 1 : (Basic Objective Questions)

1. (d)	**2.** (b)	**3.** (b)	**4.** (c)	**5.** (c)	**6.** (c)	**7.** (a)	**8.** (d)	**9.** (a)	**10.** (d)
11. (c)	**12.** (c)	**13.** (d)	**14.** (c)	**15.** (a)	**16.** (b)	**17.** (a)	**18.** (d)	**19.** (b)	**20.** (c)
21. (a)	**22.** (b)	**23.** (c)	**24.** (d)	**25.** (b)	**26.** (a)	**27.** (d)	**28.** (d)	**29.** (c)	**30.** (d)
31. (c)	**32.** (c)	**33.** (d)	**34.** (d)	**35.** (a)	**36.** (a)	**37.** (c)	**38.** (d)	**39.** (a)	**40.** (a)
41. (a)	**42.** (b)	**43.** (b)	**44.** (a)	**45.** (a)	**46.** (c)	**47.** (c)	**48.** (d)	**49.** (c)	**50.** (b)
51. (a)	**52.** (a)	**53.** (b)							

Exercise - 2 : (Previous Year JEE Mains Questions)

1. (b)	**2.** (c)	**3.** (d)	**4.** (c)	**5.** (c)	**6.** (c)	**7.** (c)	**8.** (a)	**9.** (d)	**10.** (c)
11. (c)	**12.** (c)	**13.** (d)	**14.** (b)	**15.** (d)	**16.** (a)	**17.** (c)	**18.** (c)	**19.** (c)	**20.** (b)
21. (a)	**22.** (c)								

JEE Mains Online

1. (a)	**2.** (c)	**3.** (a)	**4.** (a)	**5.** (b)	**6.** (b)	**7.** (c)	**8.** (c)	**9.** (c)	**10.** (c)
11. (c)	**12.** (c)								

Exercise - 3 : (Advanced Objective Questions)

1. (a)	**2.** (d)	**3.** (b)	**4.** (d)	**5.** (c)	**6.** (c)	**7.** (c)	**8.** (a)	**9.** (c)	**10.** (c)
11. (c)	**12.** (a)	**13.** (b,c)	**14.** (d)	**15.** (d)	**16.** (b)	**17.** (c)	**18.** (c)	**19.** (b)	**20.** (d)
21. (b)	**22.** (d)	**23.** (c)	**24.** (c)	**25.** (a)	**26.** (c)	**27.** (b)	**28.** (c)	**29.** (abcd)	**30.** (b)
31. (d)	**32.** (a)	**33.** (d)	**34.** (a)	**35.** (c)	**36.** (b)	**37.** (a,b,c)	**38.** (c)	**39.** (b)	**40.** (c)
41. (a)	**42.** (c)	**43.** (c)	**44.** (a)	**45.** (a,d)	**46.** A – S; B – P; C – Q; D – R			**47.** (b)	
48. A – Q; B – P; C – S; D – R			**49.** (c)	**50.** (a,c,d)					

Exercise - 4 : (Previous Year JEE Advanced Questions)

1. (b)	**2.** (b)	**3.** (b)	**4.** (b)	**5.** (b)	**6.** (d)	**7.** (b)	**8.** (d)	**9.** (b)	**10.** (d)
11. (c)	**12.** (a)	**13.** (a)	**14.** (b)	**15.** (c)	**16.** (c)	**17.** (a,b)	**18.** (d)	**19.** (abc)	**20.** (ab)
21. (bcd)	**22.** (c)	**23.** (c)	**24.** inert pair effect		**25.** higher effective nuclear charge				
26. electronegativity		**27.** electron affinity		**28.** (F)	**29.** (F)				